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Appendix A NUMERICAL ANSWERS TO QUESTIONS

Chemistry Review Unit

Are You Ready? (pp. 4–5)

2. (b) 4.2 g

Chapter 1

Section 1.4

Section 1.4 Questions (p. 26)

11. 1+, 2+, 3+, 3–, 2–, 1–

Chapter 2

Section 2.4

Section 2.4 Questions (p. 57)

1. (a) 18.02 g/mol
(b) 44.01 g/mol
(c) 58.44 g/mol
(d) 342.34 g/mol
(e) 252.10 g/mol

2. (a) 4 sig.dig.
(b) 2 sig.dig.
(c) 2 sig.dig.
(d) 3 sig.dig.
(e) 1 sig.dig.
(f) 4 sig.dig.

3. (a) 0.117 mol
(b) 24 g
(c) 50.0 mmol
(d) 12.49 g

4. (a) 1.72×10^3 g or 1.72 kg
(b) 50.0 L
(c) 1.55 mol/L
(d) 13.6 mL
(e) 2%
(f) 3.94 kJ

5. (a) 0.907 mol
(b) 8.56 mol
(c) 29.21 mol
(d) 1.80 mmol
(e) 2.45 mol

6. (a) 71.9 g
(b) 8.96 g
(c) 1.03 g
(d) 1.49 Mg
(e) 0.10 kg

7. (a) 2.02 g, 70.90 g, 92.92 g
(b) 64.10 g, 96.00 g, 88.02 g, 72.08 g

Unit 2

Chapter 4

Section 4.1

Lab Exercise 4.1 (p. 148)

1, 5, 3, 2, 4

Practice (p. 150)

2. (a) 0.952 atm, 724 mm Hg
(b) 110 kPa, 1.09 atm
(c) 253 kPa, 1.90×10^3 mm Hg

Practice (p. 152)

6. 263 kPa

7. 137 L

8. (b) 0.17 L.

9. 149 kPa.

Practice (p. 154)

11. –273 °C, 0 K

12. (a) $T = (0 + 273)$ K $= 273$ K
(b) $T = (100 + 272)$ K $= 373$ K
(c) $T = (-30 + 273)$ K $= 243$ K
(d) $T = (25 + 273)$ K $= 298$ K

13. (a) $t = (0 - 273)$ °C $= -273$ °C
(b) $t = (100 - 273)$ °C $= -173$ °C
(c) $t = (300 - 273)$ °C $= 27$ °C
(d) $t = (373 - 273)$ °C $= 100$ °C

Practice (p. 156)

14. (a) 16 mL

15. 0.12 L

16. 26%

17. 79 °C

Practice (p. 159)

20. 87.6 kPa

21. 7.9 L

23. 25 °C

Section 4.1 Questions (pp. 161–162)

1. (a) 8.87 kPa, 66.5 mm Hg
(b) 0.247 atm, 188 mm Hg
(c) 112 kPa, 1.11 atm

2. (a) 298 K
(b) 238 K
(c) 39 °C
(d) −65 °C

3. 384 mm Hg

4. 0.16 L

5. (a) 62 L
(b) 2.3 times larger

6. 231 °C

11. (a) 3.82:1

15. (a) 2.8 L

Section 4.2

Practice (p. 166)

5. 25.0 L

6. 0.60 L

7. (a) 1.5 ML

Section 4.2 Questions (p. 168)

4. (a) 186 kL
(b) 124 kL

5. (a) oxygen, 125 L; nitrogen monoxide, 100 L; water vapour, 150 L
(b) 375 L
(c) 33.3 L

Section 4.3

Section 4.3 Questions (p. 171)

5. 186 L

6. 2.0 mmol

7. 50.4 L

8. 0.18 mol

9. 73 mL

10. 9.80 L

11. 0.539 ML (or 539 kL)

12. 0.727 g

13. 2.58 g

14. (a) 1.8 g/L

Section 4.4

Practice (pp. 174–175)

3. 1.6 MPa

4. 41.0 mmol

5. 34 kL or 34 m^3

Lab Exercise 4.B (p. 175)

R, 8.48 kPa·L/(mol·K)

Section 4.4 Questions (p. 176)

5. 5.6 mol

6. 225 °C

8. 0.0821 $\frac{\text{atm.L}}{\text{mol·K}}$

9. (a) 34.0 g/mol

10. 22.4 L, 24.8 L

11. (a) 1.1 g/L
(d) 1.74 g/L

Unit 2 Review (pp. 180–183)

18. (a) 273 K
(b) 294 K
(c) 0 K

19. (a) 0.405 MPa
(b) 102 kPa
(c) 45.6 MPa
20. (a) 0.21 mol
(b) 0.924 mol
21. (a) 12.4 kL
(b) 1.4 ML
26. 8.23 L
27. −14 °C
28. 0.33 mol
29. 4.73 L
31. (a) 1.78 atm
(b) 196 kPa
32. 317 °C
33. (a) 302 kPa
(b) 30.7 kg
35. (a) ammonia, 1.00 L; oxygen, 1.25 L
36. (a) 50 mL
37. methane, 0.647 g/L; nitrogen, 1.13 g/L
38. (a) 150 mL
39. (b) 7.5 mmol
(c) 167 mL

Unit 3

Chapter 5

Section 5.3

Practice (pp. 205–206)

2. 7.5% V/V
3. 32% W/V
4. 4.9% W/W
5. 5.4 ppm
6. 1.8 mol/L

Practice (p. 208)

7. 350 mL
8. 7.5 kg
9. 4.1 mol
10. 0.25 mol
11. 403 mL
12. 54 mL

Practice (p. 210)

13. 15.0 g
14. 0.16 kg
(a) 355 mg
(b) 8.07 mmol/L
(a) 7.83% W/V
(b) 1.34 mol/L

Practice (p. 212)

17. (a) $[Na^+(aq)] = 0.82$ mol/L; $[S^{2-}(aq)] = 0.41$ mol/L
(b) $[Sr^{2+}(aq)] = 1.2$ mol/L; $[NO_3^-(aq)] = 2.4$ mol/L
(c) $[NH_4^+(aq)] = 0.39$ mol/L; $[PO_4^{3-}(aq)] = 0.13$ mol/L
18. $[Fe^{3+}(aq)] = 49.6$ mol/L; $[Cl^-(aq)] = 149$ mmol/L
19. (a) 11.1 g
(b) 18.5 g

Section 5.3 Questions (p. 214)

3. (a) 5% W/V
4. 20 ppm
5. (a) 0.32 mol/L
(b) $[NH_4^+(aq)] = 0.64$ mol/L; $[CO_3^{2-}(aq)] = 0.32$ mol/L
6. 79.0 g
7. 11 mg
8. 4.3 mol/L
9. (a) 0.58 g
(b) 0.75 g
(c) 1.1 g
10. 0.20 L
11. (a) 0.143 mol/L
(b) 0.429 mol/L
12. 2.3 g
17. (a) $1{:}10^9$
(b) 0.001 ppm
(c) $\frac{1 \times \text{g solute}}{\text{kg solution}}$
(d) 30 µg/kg

Section 5.4

Practice (p. 216)

1. 3.10 g
2. 200 g
4. (a) 33.2 g
5. (a) 5.93 g

Practice (pp. 218–219)

6. 42%
7. 22.5 mL
8. (a) 0.250 mmol/L
(b) 0.399 mg

Section 5.4 Questions (p. 219)

3. 3.27 g
4. 43.5 L
5. 6.85 g/L
6. 1.51 g
7. 25.0 mL
8. (a) 1.30 g
(b) initial volume = 10.0 mL

Chapter 5 Review (pp. 231–233)

22. 51 g, 150 g, 250 g
23. 0.3 L
24. (a) 0.70 mol/L
(b) 0.125 mol/L
(c) 2.0 mol/L
(d) 0.66 mmol/L
25. 12.6 g
26. 42.8 mL
27. 6.6 g
28. (a) 56 mg
32. 28.1 mL
34. (a) $[K^+(aq)] = 0.14$ mol/L; $[NO_3^-(aq)] = 0.14$ mol/L
(b) $[Ca^{2+}(aq)] = 0.14$ mol/L; $[Cl^-(aq)] = 0.28$ mol/L
(c) $[NH_4^+(aq)] = 0.42$ mol/L; $[PO_4^{3-}(aq)] = 0.14$ mol/L

Chapter 6

Section 6.2

Practice (p. 239)

1. (a) 10^{-7} mol/L
(b) 10^{-11} mol/L
(c) 10^{-2} mol/L
(d) 10^{-4} mol/L
(e) 10^{-14} mol/L
2. (a) 3
(b) 5
(c) 7
(d) 10
3. 100

Practice (p. 242)

5. (a) 2.68
(b) 5.0
(c) 6.602
(d) 8.14
6. (a) 5×10^{-12} mol/L
(b) 2.2×10^{-3} mol/L
(c) 6×10^{-5} mol/L
(d) 1.76×10^{-14} mol/L

Practice (p. 243)

9. −0.65
10. 1.6×10^{-7} mol/L

Section 6.2 Questions (p. 244)

6. 5.00

8. increase of 3 pH units

9. (a) −1.00

(b) 14.80

Section 6.3

Lab Exercise 6.B (p. 247)

4.4 - 4.8; 6.0 - 6.6; 3.2 - 3.8

Section 6.3 Questions (p. 247)

2. (a) < 4.8

(b) > 12.0

(c) > 5.4

(d) between 6.0 and 7.6

3. (a) between 5.4 and 6.0

(b) 1×10^{-6} mol/L

Section 6.5

Practice (p. 255)

4. (a) 0.15 mol/L

(b) 0.82

Chapter 6 Review (pp. 263–264)

14. (a) 2.74×10^{-12} mol/L

(b) 3×10^{-4} mol/L

15. 10×

17. (a) between 5.4 and 6.0

(b) 1×10^{-6} mol/L to 4×10^{-6} mol/L

Unit 3 Review (pp. 265–269)

22. (a) 75.0 ppm

(b) 1.87 mmol/L

23. 1.2 L

24. (a) 0.23 mol/L

(b) 20 mmol

(c) 0.103 L

(d) 0.155 mol/L

(e) 1.21 g

25. 16.9 mg/L

26. (a) 14 L

27. (a) $[Na^+(aq)] = 4.48$ mol/L; $[S^{2-}(aq)] = 2.24$ mol/L

(b) $[Fe^{2+}(aq)] = 0.44$ mol/L; $[NO_3^-(aq)] = 0.88$ mol/L

(c) $[K^+(aq)] = 0.525$ mol/L; $[PO_4^{3-}(aq)] = 0.175$ mol/L

(d) $[Co^{3+}(aq)] = 0.0862$ mol/L; $[SO_4^{2-}(aq)] = 0.129$ mol/L

30. (a) 2.12

(b) 2.60

31. (a) 2.74×10^{-12} mol/L

(b) 3×10^{-4} mol/L

35. (a) 10^{-3} mol/L (fruit juice); 10^{-12} mol/L (cleaning solution)

(b) 10^9:1

37. (a) 6.3

(b) 5.10

39. 2.82 g

49. 0.012 mol/L

51. 0

56. (b) 4×10^{-3} mol/L

Unit 4

Are You Ready? (pp. 272–273)

3. (a) 2×10^{-5} mol/L

4. $(NH_4)_3PO4(s)$, 0.2951 mol; $CH_3COOH(l)$, 3.5 g

5. NaOH(aq), 1.10 mol; HCl(aq), 0.00345 L; $Na_2SO_4(aq)$, 1.13 mol/L

6. $CH_4(g)$, 0.611; $UF_6(g)$, 0.120; $CO_2(g)$, 2.48×10^3 L; Ar(g) 0.161

Chapter 7

Section 7.2

Practice (p. 290)

9. 12 g

10. 66.2 g

11. 36.3 g

12. 6.11 g

13. 0.307 g

14. 2.32 g

Lab Exercise 7.A (p. 291)

6.68 g

Lab Exercise 7.B (p. 293)

9.12 g

Section 7.2 Questions (p. 293)

6. 3.88 g

8. (a) 2.93 g

(b) 95.9%

10. (a) 11 g

(b) 91%

Section 7.3

Practice (p. 296)

1. 16 L

2. 60.1 kL

3. 150 L

Section 7.3 Questions (p. 298–299)

2. 6.0 L

3. 2.62 ML

4. 232 kg

5. 0.21 kL, or 0.21 m^3

6. 0.77 L

7. 1.94 kg

8. 0.57 L

Section 7.4

Practice (p. 302)

1. 0.537 mol/L

2. 375 mL

3. 90.0 mL

Lab Exercise 7.C (p. 302)

210 mg; 0.20 g

Lab Exercise 7.D (p. 303)

0.351 mol/L

Section 7.4 Questions (p. 303)

1. 0.35 L

2. 17.8 mol/L

3. 23.9 mmol/L

4. (b) 624 mg or 0.624 g

(c) 98.0%

5. 639 g

Chapter 7 Review (p. 309–311)

20. (a) $NaHCO_3(s)$, 2.4 mol; $Na_2CO_3(s)$, 1.2 mol; $CO_2(g)$, 1.2 mol; $H_2O(l)$, 1.2 mol

(b) 0.631 kg

21. 46 mL

2.93 g

23. (a) 95.0%

26. (a) 14.7 g

(b) 10.7 g

28. 2.08 kg

29. 1.91 kg

30. 1.17 kg

33. 0.668 mol/L

Chapter 8

Section 8.2

Lab Exercise 8.A (p. 317)

5.40 g

Section 8.2 Questions (p. 319)

1. 97 g sodium sulfate

2. 0.135 mol/L

3. 0.210 g predicted; 0.20 g obtained; 5%

4. 0.351 mol/L

Section 8.3

Practice (p. 321)

1. (a) 1.1 g to 1.2 g

2. 33.6 mL

Practice (p. 324)

3. (a) 5.0 mol

(b) 0.55 mol

(c) 0.26 mol

(d) 5.46 mmol

4. (a) 6.25 g
(b) 2.86 g
(c) 36.mg
(d) 4.2 kg

Section 8.3 Questions (p. 327)

4. 3 g
5. (a) 2.77 g
(b) 3.89 g
6. 97.9%
7. (b) 2.3 g $BaCl_2$
(c) 8.4 g
8. (b) 3.5 g Zn
(c) 83%
9. (b) 125 mL
(c) 138 mL

Section 8.4

Section 8.4 Questions (p. 332)

1. $[NH_3] = 20.7$ mol/L
4. 0.660 mmol

Section 8.5

Web Activity: Web Quest—Blood Alcohol Content (p. 333)

Analysis
(a) 3.00×10^{-5} mol
(b) 2.22×10^{-6} mol
(c) 2.778×10^{-5} mol
(d) 0.0139 mol/L; BAC = 0.064 g/100 mL

Practice (p. 336)

2. (b) 7

Section 8.5 Questions (p. 339)

8. (a) 1.27 g
9. 0.140 mol/L

Chapter 8 Review (pp. 346–348)

13. (a) 0.075 mol lead (II) nitrate
(b) 1.00 mol propane
(c) 0.50 mol zinc
(d) 50 mmol sulfuric acid
14. 79.7%
15. 52.4 mL
16. 2.70 g
17. (a) 20 mL, 20 mL, 20 mL, 20 mL
(b) 7, 9, 5, 7
19. 21.3 mmol/L

Unit 4 Review (pp. 349–51)

24. (a) 106%
25. 699 kg
26. (a) 0.20 mol Zn(s)
(b) 5.0 mmol Cl_2(aq)
(c) 0.05 mol NaOH(aq)
27. 0.12 kg
31. 13.0 mL
33. 2.93 g
34. (b) 1.07 g
36. 1.22 g predicted; 1.27 g obtained; 4%

Unit 5

Chapter 10

Section 10.3

Practice (p. 431)

15. (a) 18 mL
(b) 18 mL
(c) 18 mL

Chapter 10 Review (pp. 466–467)

20. 36.8%

Unit 5 Review (pp. 119–120)

40. (a) 0.46 kL
(b) 4.6×10^3 km
42. 78.2%
48. 96.3%; 94.8%

Unit 6

Chapter 11

Section 11.1

Case Study, (p. 482)

2. 77.8 kJ

Section 11.1 Questions (pp. 483–484)

1. (a) respectively, for 2003: 8.22%, 35.21%, 43.55%, 3.99%, 9.04%
2. (a) respectively, for 2003: 58%, 63%, 56%, 42%, 8%
3. respectively, for 2003: 30.4%, 29.5%, 2.8%, 17.7%, 1.7%, 17.9%

Section 11.2

Practice (p. 487)

6. 84 kJ
7. 0.39 MJ
8. 67.7%

Practice (p. 492)

12. −5.00 MJ
13. −612 kJ

Lab Exercise 11.A (p. 492)

−54.6 kJ/mol

Section 11.2 Questions (p. 494)

2. (a) 2.1 kJ
(b) 2.2 kJ
(c) 4%
3. (a) −7.8 MJ
(b) −2.07 MJ
4. 1.54 g
5. −0.242 MJ
6. −64 kJ/mol
7. −26 kJ/mol
11. −41.1 kJ/mol

Section 11.3

Section 11.3 Questions (p. 501)

3. (a) 241.8 kJ/mol
(b) −318.0 kJ/mol
(c) +81.6 kJ/mol
(d) −372.8 kJ/mol
4. (a) −114 kJ
(c) −114 kJ/mol
(d) −57 kJ/mol

Section 11.4

Practice in 11.4 (pp. 504–505)

1. −851.5 kJ
2. +131.3 kJ
3. −524.8 kJ
4. −205.7 kJ

Lab Exercise 11.B (p. 506)

−3488.7 kJ/mol; -3.35 MJ/mol; 4.0%

Lab Exercise 11.C (p. 506)

+205.9 kJ

Section 11.4 Questions (pp. 508–509)

1. −69.5 kJ
2. (a) +136.4 kJ
(b) −44.2 kJ
3. +121.2 kJ/mol
7. −2.39 MJ
8. −2.20 MJ
9. −250.1 kJ/mol
10. −492 kJ
11. −21.8 kJ

Section 11.5

Lab Exercise 11.D (p. 513)

−725.9 kJ/mol; −598 kJ/mol; 17.6%

Section 11.5 Questions (pp. 514–515)

2. (a) +205.9 kJ
(b) −41.2 kJ
(c) −91.8 kJ

3. (a) −225.5 kJ/mol
(b) −58.1 kJ/mol
(c) −23.6 kJ/mol
4. (a) −145.6 kJ/mol
5. (a) −349.6 kJ/mol
6. (a) −562.0 kJ
8. (a) +179.2 kJ
(b) −64.5 kJ
(c) +114.7 kJ
(d) +114.7 kJ
9. –114.7 kJ
10. −198.7 kJ/mol
12. −2803.1 kJ/mol; −2.80 MJ/mol; 0.143%

Chapter 11 Review (pp. 519–521)

17. (a) 44 kJ
(b) 36 kJ/g
18. (a) 75.1 kJ
(b) −75.1 kJ
(c) −2.87 MJ/mol
19. −44.4 kJ/mol
20. (a) 21.8 MJ
(b) 0.436 kg
(c) 234 L
22. −1 256.2 kJ
24. −107.4 kJ
25. −69 kJ
26. (a) −401.0 kJ/mol
27. (a) −27.2 kJ
(b) +2.3 kJ
28. −426 kJ; −110 kJ; −602 kJ
29. (a) −2 219.9 kJ/mol
(b) −2 043.9 kJ/mol
(c) −1 564.2 kJ/mol

Chapter 12

Section 12.3 Questions (p. 70)

3. (a) +60 kJ
(b) +95 kJ
(c) −35 kJ
(d) +35 kJ
7. 5 kJ/mol; 6 kJ/mol; 11%

Chapter 12 Review (pp. 545–546)

16. (a) 52 kJ
(b) 90 kJ
(c) 22 kJ
(d) 60 kJ
(e) −38 kJ
(f) 38 kJ
(g) −38 kJ
(h) 38 kJ

Unit 6 Review (pp. 547–551)

19. (a) −176.2 kJ
20. (b) −128.7 kJ
(c) −4.02 MJ
21. (a) −1164.8 kJ
(b) −388.3 kJ/mol
(c) −1.51 GJ
22. (a) +136.4 kJ
(b) +311.4 kJ
23. (a) −23.5 kJ
28. (a) −98.0 kJ/mol
29. −890.5 kJ/mol; −1560.4 kJ/mol; −2219.9 kJ/mol
−55.48 MJ/kg; 51.88 MJ/kg; 50.33 MJ/kg
62.31 mol/kg; 66.49 mol/kg; 68.01 mol/kg
30. (b) −134.6 kJ
31. (a) −1366.8 kJ
36. −25.6 kJ/mol

Unit 7

Chapter 13

Section 13.3

Practice (p. 585)

1. (a) +4
(b) +7
(c) +6
(d) +6
(e) –1
(f) –1
2. (a) +1
(b) +2
(c) +4
(d) –3
(e) –2
(f) +5
(g) 0
(h) –3
3. (a) 0
(b) 0
(c) +4
(d) +2
5. –4; –2; 0; +2; +4

Section 13.3 Questions (p. 595)

4. (a) C +2; O –2
(b) O 0
(c) N −3; H +1; Cl −1
(d) H +1; P +5; O −2
(e) Na +1; S +2; O −2
(f) Na +1; P +5; O −2
5. (a) $+\frac{8}{3}$

Section 13.4

Lab Exercise 13.C (p. 598)

0.258 mol/L

Practice (p. 598)

3. 8.5 L
4. 0.325 mol/L
5. 11.4 mmol/L

Lab Exercise 13.D (p. 599)

1.31 mol/L

Investigation 13.4 (p. 603)

Analysis: 2.94%

Section 13.4 Questions (p. 600)

4. 1.92 mol/L
5. Analysis: 74.9 mmol/L
Evaluation: 6.4%
6. 25.9 mmol/L
7. 29.8 mmol/L

Chapter 13 Review (pp. 606–609)

29. 4.8 g
30. 56.3 mmol/L
33. Analysis: –33 °C

Chapter 14

Section 14.1

Practice (p. 614)

3. 6

Section 14.2

Practice (p. 631)

11. (a) +0.77 V
(b) +0.45 V
(c) +1.23 V

Practice (p. 633)

12. (a) +0.47 V
(b) +0.50 V
(c) +0.77 V
13. −0.34 V
14. Cu: +3.38 V; Zn: +2.28 V

Lab Exercise 14.A (p. 633)

$Cr_2O_7^{2-}$: +1.23 V; Pd^{2+}: +0.95 V; Tl^{+}: −0.34 V; Ti^{2+}: −1.63 V

Section 14.2 Questions (pp. 637–638)

10. (a) +0.48 V

12. −0.28 V

15. +1.56 V

Section 14.3

Practice (p. 644)

5. (a) +0.80 V
(b) +1.23 V

6. (a) +1.48 V
(b) 0.00 V

Practice (p. 645)

12. (a) +0.43 V
(b) +0.29 V

13. −1.30 V

Section 14.4

Practice (p. 653)

1. 45 C

2. 3.89 A

3. 7.13 C

4. 3.91 min

Practice (p. 654)

5. 41 mmol

6. 3.16 h

7. 1.2 A

Section 14.4 Questions (p. 657)

1. 2.80 mmol

2. 0.58 t

3. 82.8 min

4. 52.8 kA

5. (a) 1.63 t
(b) 4.76 t

6. 0.174 mol/L

7. 24.42 g

8. 0.102 g

9. (a) 26 g
(b) 25.4 min

10. 21. 3 g

11. 25.0 A

12. 0.766 g; 0.71 g; 7%

14. (c) 483 g
(d) 41.9 g

Chapter 14 Review (p. 665)

18. (d) +0.93 V

21. (d) 0.00 V

23. 50.3 kg

24. 0.125 A

Unit 7 Review (pp. 667–669)

41. 5.78 g

42. 8.56 mmol/L

45. (a) 1.22 V
(b) 0.80
(c) 0.00 V

46. (a) 1.90 V
(b) 1.23 V
(c) 1.64 V

47. (b) 1.23 V

48. (b) 2.19 V

49. (a) 1.99 V
(b) 590 s (or 9.84 min)

50. 2.98 kA

51. 20.1 min

52. 1.0 kmol/h

54. (b) +1.20 V

Unit 8

Are You Ready? (pp. 672–673)

5. (a) 5.94 mol/L
(b) 0.220 mol/L
(c) 0.0403 mol/L
(d) 0.0446 mol/L

Chapter 15

Section 15.1

Practice (p. 682)

4. (a) 2.00 mol
(b) 70.0%

6. (c) 1.00 L
(d) 60%

Lab Exercise 15.B (p. 686)

0.46

Section 15.1 Questions (pp. 688–689)

3. (b) 14 mol

5. 51

6. 0.11 mol/L

7. (c) 0.200 mol
(d) 0.80 mol
(e) 0.40 mol
(f) $[HBr(g)] = 0.100$ mol/L; $[H_2(g)] = 0.20$ mol/L; $[Br_2(g)] = 0.20$ mol/L
(g) 4.0

8. 54.1

9. 1.5

10. (a) 0.78 mol/L
(b) 0.39 mol

Chapter 15 Review (p. 705–706)

19. 0.0013 mol/ L

20. 0.0403 mol/ L

21. 0.032

26. (a) 0.068

27. (a) 5.07 mol/L
(b) 7.60 mol/L

28. 1.00 mol/L

32. (b) 0.0029 mol/L

Chapter 16

Section 16.1

Practice (p. 716)

1. 2.3×10^{-12} mol/L

2. 3.3×10^{-11} mol/L

3. 2.5×10^{-13} mol/L

4. 7.2×10^{-13} mol/L

5. 1.4×10^{-14} mol/L

6. 1.8×10^{-7} %

Practice (p. 718)

7. (a) 1.8×10^{-12}; 2.26; 11.74
4×10^{-9}; 3×10^{-6}; 8.4
5.0×10^{-4}; 3.30; 10.70
4.0×10^{-4}; 2.5×10^{-11}; 3.40

8. 14.64; −0.64

9. 0.09 g

Section 16.1 Questions (p. 721)

4. (b) 7.8×10^{-6} mol/L; 5.11

5. 7.7×10^{-12} mol/L

6. 7.40

7. 3×10^{-6} mol/L

8. 1000 (10^3)

9. 0.372

10. 0.25 g

11. 4.27

12. −0.016; 14.02

13. 2.42; 11.58

14. 18 mg

Section 16.3

Practice (p. 743)

2. (a) 0.20 mol/L
(c) 1.9×10^{-3} mol/L
(d) 2.64
(f) 4.84

5. (a) 1.16%
(b) 1.36×10^{-5}

6. (a) 3.7%
(b) 1.35×10^{-5}; 1.4×10^{-4}

7. (b) 3.5×10^{-2} mol/L; 3.5×10^{-2} mol/L; 1.45; 1.8%

8. (a) 0.26 mol/L; 0.58; 2.6%

9. (a) 8.2×10^{-5}

Practice (p. 746)

12. 7.7×10^{-10}

13. 4.2×10^{-10}

Section 16.3 Questions (p. 750)

1. 10.27

2. 11.23

4. 4.77

5. 1.4×10^{-3}

6. 1.3×10^{-10}

7. 4.2×10^{-4} mol/L

10. (a) 9.95
(b) 4.25
(c) 3.85

Section 16.4

Practice (pp. 762–763)

13. (a) <7
(b) >7
(c) ~ 7
(d) <7

Chapter 16 Review (p. 773)

17. (a) 0.10 mol/L; 1.00
(c) 9.4×10^{-5} mol/L; 4.03

18. (a) 5.8×10^{-7}

20. 7; >7

21. 10.0 to 11.4; 2.5×10^{-3} to 2.5×10^{-3} mol/L

22. 0.019%; 8.46; 5.54; 2.9×10^{-6} mol/L

Unit 8 Review (pp. 775–779)

26. 0.62 mol/L; 0.62 mol/L; 1.38 mol/L; 1.38 mol/L

29. 5.0 to 5.2; 1×10^{-5} mol/L to 6×10^{-6} mol/L

35. (c) 25 mL; 50 mL

36. 3.48 mol/L; 1.04 mol/L

37. 2.3

38. 0.882 mol/L

39. (a) (i) 7.00; 10 mL
(ii) 13.00
(iii) 12.60
(v) 11.8
(vi) 1.9

45. 1.24%; 1.6×10^{-5}

46. 6.5×10^{-3} %; 4.2×10^{-10}

47. 1.5×10^{-4} mol/L; 0.58; 0.30%

48. (b) 3.1×10^{-2} mol/L; 1.51; 1.5%
(c) 2.1×10^{-11} mol/L

49. 0.012%; 8.14; 5.86; 1.4×10^{-6} mol/L

50. 7.9×10^{-10}

51. (a) 1.5×10^{-3}; 5.0×10^{-2}; 1.9×10^{-1}

B.1 Scientific Problem-Solving Model

Scientists ask questions and seek concepts to answer these questions by applying consistent, logical reasoning to describe, explain, and predict observations, and by doing experiments to test hypotheses or predictions from these hypotheses. In this way science progresses using a general model for solving problems and employing specific processes as part of a problem-solving strategy.

Every investigation in science has a *purpose*; for example:

- to *create* a scientific concept (a theory, law, generalization, or definition)
- to *test* a scientific concept
- to *use* a scientific concept, e.g., in chemical analysis

Once you know the purpose, you need a problem and a general design. For example, if the purpose is to perform a chemical analysis to determine the quantity of a substance, then possible designs include distillation and precipitation. Once you choose a design, there are many specific questions that you might ask, many possible reactants you might choose, and many other variables you might need to consider.

B.2 Investigation Report Outline

An investigation report is the final result of your problem solving. Your report should follow the model outlined in **Figure 1**. As a further guide, use the information and instructions for the specific processes listed below. The parts of the investigation report that you are to provide are indicated in the text in a checklist (**Figure 2**).

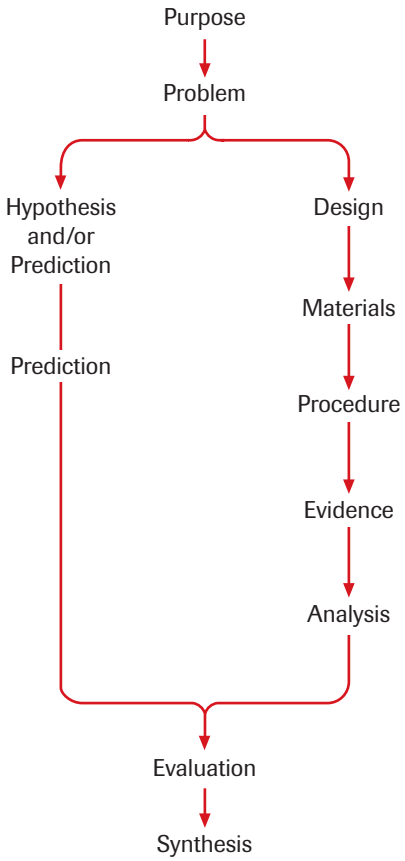


Figure 1
A scientific problem-solving model helps to guide your laboratory work, but does not illustrate the complexity of the work.

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	● Evaluation (1, 2 and/or 3)
○ Hypothesis	○ Procedure	
○ Prediction	● Evidence	

Figure 2
Shaded circles indicate the parts you are expected to complete in a particular investigation report. One or more parts of an Evaluation may be required, as indicated by the numbers.

Purpose

Although this is usually provided, you will be expected to identify the purpose of an investigation before, during, and after your laboratory work. Most often, the purpose is to create, test, or use a chemistry concept.

Problem

The Problem is a specific question to be answered in the investigation. If appropriate, you should state the question in terms of manipulated and responding variables. In most cases, the problem is chosen for you. Only when creating a concept will the Purpose and the Problem be the same.

Hypothesis

The hypothesis is an (often untested) empirical or theoretical concept that provides a possible explanation for a natural or technological phenomenon. Only some kinds of investigations require a hypothesis, such as investigations that test a hypothesis using a general question as the Problem.

Prediction

The Prediction is the expected answer to the Problem according to a scientific concept (for example, a hypothesis, theory, law, or generalization) or another authority (for

example, a reference source or a label on a bottle). Write your Prediction using the format, "According to [an authority], [answer to the Problem]." Include your qualitative and quantitative reasoning (based upon the authority) with the Prediction.

Design

The Design provides a brief overview of the Procedure to obtain an answer to the Problem. Included in the Design are reacting chemicals and, if applicable, brief descriptions of diagnostic tests, variables, and controls. Write your Design as a paragraph of one to three sentences.

Materials

This section consists of a complete list of all equipment and chemicals, in columns, including sizes and quantities. Appendix C.3, page 797, shows laboratory equipment, including common sizes.

Procedure

The Procedure is a detailed set of instructions designed to obtain the evidence needed to answer the Problem. Write a list of numbered steps in the correct sequence, including any safety and waste disposal instructions (Appendix E, page 807). Whenever possible, repeat measurements several times.

Evidence

The Evidence includes all qualitative and quantitative observations relevant to answering the Problem. Organize your evidence in tables whenever possible (Appendix F.4, page 815). Be as precise as possible in your measurements and include any unexpected observations that may affect your answer and its certainty (in significant digits). Scientific honesty demands that you report all evidence collected and not just the evidence you think is correct or "normal."

Analysis

The Analysis includes manipulations, interpretations, and calculations based on the evidence. Tables and graphs that include or facilitate interpretations and calculations are included in the Analysis. You may need to differentiate between relevant and irrelevant observations. Communicate your work clearly and logically. Conclude the Analysis with a statement of your experimental answer to the Problem, including a phrase such as, "According to the evidence gathered in this experiment, [answer to the Problem statement]."

Evaluation

Part 1 of the Evaluation of an investigation that you actually perform usually involves judging the validity of the experiment (Design, Materials, Procedure, and Skills). Evaluation of Lab Exercises (simulated investigations) will not usually involve Part 1. Only if you are confident that no major flaws are present can you proceed to the second part. In Part 2, you use the results of the experiment to evaluate the Prediction (if one was made) and the Hypothesis (if there is one). This assumes that a prediction is being tested in an experiment. If it is the experimental design that is being tested, then the two parts of the evaluation would be reversed and the percent difference is used to judge the success of the design. The last part of the Evaluation, Part 3, comes back, full circle, to the Purpose of the investigation. Was the Purpose fulfilled?

The parts of the Evaluation you will be expected to complete are shown in parentheses after Evaluation in the Report Checklist.

Write your Evaluation in paragraph form, using the topic sentences suggested below or an adaptation of them. Some of the more important criteria for a judgment are listed as questions; use selected questions to guide your judgments. Show as much independent, critical, and creative thought as possible in support of your judgments.

Part 1. Evaluation of the Investigation

- "The design of the investigation [name or describe in a few words] is judged to be adequate/inadequate because …"

  Were you able to answer the Problem using the chosen experimental design? Are there any obvious flaws in the design? What alternative designs (better or worse) are available? As far as you know, is this design the best available in terms of controls, variables, efficiency, and safety?

- "The materials are judged to be adequate/inadequate because ..."

  Did you have all of the necessary materials? Was the equipment of reasonable quality? What materials could be improved to obtain better results?

- "The procedure is judged to be adequate/inadequate because …"

  Were the steps that you used in the laboratory correctly sequenced, and adequate to gather sufficient evidence? What improvements could be made to the procedure, such as more trials?

- "The technological skills are judged to be adequate/inadequate because …"

  Which specialized skills, if any, might have the greatest effect on the evidence gathered? Was the evidence from repeated trials reasonably similar?

How can the evidence gathered be improved?

- "Based upon my evaluation of the experiment, I am not/moderately/very certain of my experimental evidence. The sources of uncertainty or error are …"

  State the sources and your confidence in the experimental evidence. What would be an acceptable total of the experimental error (1%, 5%, or 10%)?

Part 2. Evaluation of the Prediction and Authority Being Tested

- If applicable, "the percent difference between the experimental result and the predicted value is…"

  $$\% \text{ difference} = \frac{|\text{experimental value} - \text{predicted value}|}{|\text{predicted value}|} \times 100$$

  How does this difference compare with your estimated total uncertainty or experimental error?

- "The prediction is judged to be verified/inconclusive/falsified because …"

  Does the predicted answer clearly agree with the experimental answer reported in your analysis? Can the percent difference be accounted for by the sources of error (percent error) listed above?

- "The authority being tested [name the authority or hypothesis (reasoning)] is judged to be acceptable/unacceptable in this experiment because …"

  Was the prediction verified, inconclusive, or falsified? How confident do you feel about your judgment?

Part 3. Revisiting the Purpose

Did you accomplish the Purpose of this investigation? Is there a need for additional investigations to better achieve the Purpose?

Notes on Data and Evidence

Data

Data may be found on data sheets, in data tables, and in databases. Data from one of these sources can become evidence with the purpose to create, test, or use a scientific concept. Evidence is data with a scientific purpose.

Common Sources of Experimental Error

- conditions (e.g., SATP) not controlled
- impure reactants or products
- any measurement process
- incomplete reaction
- judgment of colour (e.g., indicator)
- loss of solid in a filtration (stuck to glass or passed through filter)
- incomplete drying of a product
- manipulative skill

Do not use "human error" as a source of uncertainly or experimental error.

Percent Yield versus Percent Difference

In most experiments, a percent difference is a measure of accuracy. In some experiments in which a product is collected and measured, a percent yield is used instead of a percent difference.

$$\% \text{ yield} = \frac{\text{actual quantity obtained}}{\text{predicted (maximum) quantity} \times 100}$$

Experimental Error and Percent Difference

Some people and books use the term "percent error" in place of "percent difference." In this textbook, we use two percentages: One is an estimate of the total expected *experimental error* (Evaluation, Part 1), and the other is the actual difference (the *percent difference*) you determined based on your prediction and analysis (Evaluation, Part 2). The crucial point is how these two percentages compare. No experiment can ever be expected to be perfect. For example, if the equipment you used is only precise to +/− 5%, then any percent difference you obtain that is less than or equal to 5% is as good a result as can be expected. If the percent difference is larger than the reasonably acceptable experimental error, then the prediction is falsified.

If the percent difference is equal to or less than the experimental error, then the prediction is verified.

Replication

An authority may be judged unacceptable in one experiment. This does not mean the authority is immediately discarded. Replication by independent workers is always required to refute any accepted theory.

B.3 Sample Investigation Report

The Reaction of Hydrochloric Acid with Zinc

Report Checklist

○ Purpose	● Design	● Analysis
○ Problem	● Materials	● Evaluation (1, 2, 3)
○ Hypothesis	● Procedure	
● Prediction	● Evidence	

Purpose

The purpose of this investigation is to test one of the ideas of the collision–reaction theory.

Problem

How does changing the concentration of hydrochloric acid affect the time required for the reaction of hydrochloric acid with zinc?

Prediction

According to the collision–reaction theory, if the concentration of hydrochloric acid is increased, then the time required for the reaction with zinc will decrease. The reasoning that supports the prediction is that a higher concentration produces more collisions per second between the hydrochloric acid entities and the zinc atoms. More collisions per second would produce more reactions per second and, therefore, a shorter time required to consume the zinc.

Design

Different known concentrations of excess hydrochloric acid react with the same quantity of zinc metal. The time for the zinc to completely react is measured for each concentration of acid solution. The concentration of hydrochloric acid is the manipulated variable and time is the responding variable. The temperature, mass, and surface area of zinc, and volume of acid are the controlled variables.

Materials

lab apron
eye protection
(4) 10 mL graduated cylinders
(4) 18 × 150 mm test tubes and test-tube rack
clock or watch (precise to the nearest second)
four pieces of a zinc metal strip (5 mm × 5 mm)
HCl(aq): 2.0 mol/L, 1.5 mol/L, 1.0 mol/L, 0.5 mol/L
a weak base (e.g., baking soda)

Procedure

1. Transfer 10 mL of 2.0 mol/L HCl(aq) into an 18 × 150 mm test tube. Avoid contact with skin, eyes, clothing, or the desk. If you spill this acid on your skin, wash immediately with lots of cool water.
2. Carefully place a piece of Zn(s) into the hydrochloric acid solution and note the starting time of the reaction.
3. Measure and record the time required for all of the zinc to react.
4. Repeat steps 1 to 3 using 1.5 mol/L, 1.0 mol/L, and 0.5 mol/L HCl(aq).
5. Neutralize the acid with a weak base and then pour it down the sink with the water running.

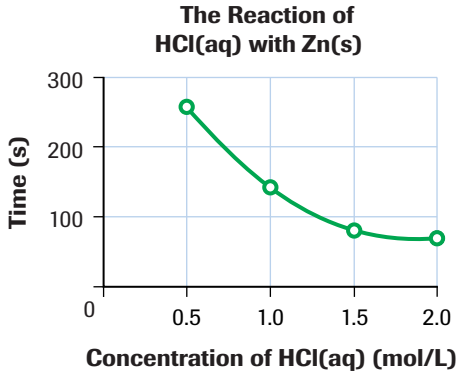
Evidence

Gas bubbles formed immediately on the surface of the zinc strip when it was placed into the hydrochloric acid solution. The bubbles appeared to form more rapidly when the concentration of the acid was higher.

The Reaction of HCl(aq) with Zn(s)

Concentration of HCl(aq) (mol/L)	Time for reaction (s)
2.0	70
1.5	80
1.0	144
0.5	258

Analysis



According to the evidence obtained, increasing the concentration of hydrochloric acid decreases the time required for the complete reaction of a fixed quantity of zinc.

Evaluation

The design, reacting zinc with excess hydrochloric acid, is judged to be adequate because this experiment produced the type of evidence needed to answer the problem with a high degree of certainty. There are no obvious flaws in this design. In my judgment, the design is efficient and safe, and all necessary variables are controlled.

The materials appear to be adequate because the quality of the evidence was sufficient to give a clear answer to the problem. Using a pipette would provide better precision than a graduated cylinder but this would not change the overall result.

The procedure is also judged adequate because it produced sufficient evidence. Possible improvements include extending the range of concentrations and performing more trials for each concentration. The reaction could be done in a container that allowed better mixing.

The technological skills are judged adequate because no specialized skills were involved. Timing the start of the reaction may have some uncertainty but the lack of multiple trials makes this difficult to judge.

Based upon my evaluation of the experiment, I am very certain about the experimental results. Sources of uncertainty in this investigation include: measurement errors for volume and time, the purity and surface area of the zinc metal strip, the concentration of the acid, and a little uncertainty in estimating when the last bit of zinc had reacted. An estimate of the total effect of all experimental uncertainties is about 5%.

The prediction is judged to be verified because the qualitative observations and the graph clearly indicate that the reaction time decreases as the concentration increases. There is little deviation from a smooth curve in the graphed results.

The collision–reaction theory is judged to be acceptable in this experiment because the prediction was clearly verified. I am quite confident in this judgment because other groups in the class obtained similar results.

The purpose of this investigation—to test one idea of the collision–reaction theory—was accomplished but only for one reaction. Replication with many other reactions would need to be investigated to have a more valid test.

B.4 The Nature of Scientific Research

Citizens in a democratic society are often required to read and interpret media reports of scientific research. Health and environment research reports are, for example, commonly portrayed in the media. Sometimes the research reports appear to contradict each other and sometimes the reports promote more uncertainty than certainty. Understanding the terminology and concepts for describing a research study is increasingly important for responsible citizenry. Listed below are some of the terms and concepts that will help you both answer questions in this textbook and understand and critique media reports of research.

Types of Studies

correlational study—the connection or degree of agreement (e.g., –0.3, +1.0) is sought between two variables, often without controlling for other variables; correlational studies often lead to cause-and-effect studies

cause-and-effect study—one variable is manipulated and all other variables, other than the responding variable, are controlled

control experiment—see cause-and-effect study

clinical trial—a controlled study involving people; usually a final-stage, double-blind study

Design Factors

term of study—the duration of the experiment e.g., observations over 5 s, 30 min, 3 mon or 15 a; long-term studies are most often preferred

sample size—the number of entities or people in a study; generally large sample sizes are preferred

random sample—one chosen randomly from the population of entities (to reduce bias)

replication—repetition of a study, generally, by an independent research group

placebo—in medicinal experiments, an inactive item (e.g., sugar pill) or treatment given to the control group

placebo effect—a beneficial effect arising from a patient's expectations; present in both the control group and the experimental group

single blind—the subject (e.g., patient) does not know whether she/he has received the treatment or a placebo, but the experimenter knows

double blind—neither the subject (e.g., patient) nor the directly involved experimenter knows whether the subject has received the treatment or a placebo

control—a standard or comparison value, or procedure (e.g., leaving one of several identical samples unaltered for comparison), or a placebo

control group—a comparison group that does not receive the experimental treatment
experimental group—a group that receives the experimental treatment

Nature of Evidence and Results

anecdotal—based upon personal experience or hearsay
reliable—reproducible or consistent time after time
valid—judged to be supported by adequate designs, materials, procedures, and skills
accurate—judged to be true or agreeing with the accepted value
precise—closely related or very similar; related to reproducibility of results
statistical bias—a sampling or testing error caused by systematically favouring some outcomes over others
random result—a result that could be expected on the basis of probability (e.g., 50% heads and tails when flipping a large number of coins)
coincidence—a result that is accidental and irrelevant to the study
significant difference—a difference that is greater than could be randomly expected when an experimental group and a control group are compared
certainty—the degree to which something is accepted by an individual or community (e.g., the evidence may have a high or low degree of certainty); measured by, for example, counting significant digits

Reporting Research

refereed journal—an academic journal for which research papers are sent to subject experts to determine whether the report is of sufficient quality to publish; also called *peer-reviewed journal*
abstract—a short summary describing the research processes and results

Science–Technology–Society (STS) Issues

risk–benefit analysis—a process of gathering and analyzing evidence that leads to decision making (and to an evaluation of the process itself)
stewardship—actively supervising and managing an entity or event (e.g, the environment)
perspective—a point of view or way of analyzing an object or event
multiperspective—based upon positive and negative evidence and arguments from many perspectives (e.g., scientific, technological, economic, environmental, political, legal, ethical, social, and emotional)
ways of knowing—methods used to obtain knowledge or information; examples include traditional (Aboriginal), empirical, theoretical, referenced, and memorized

Scientific Attitudes

scientific attitude—a disposition or demonstration of feelings or thoughts (e.g., honesty, objectivity, willingness to change, respect for evidence, critical mindedness, suspended judgment, open-mindedness, and questioning predisposition)
tolerance of uncertainty—the degree to which people and institutions tolerate uncertainty (without claiming absolute certainty), although they strive for greater and greater certainty

COMMUNICATION example 1

Create an experimental design to test a new drug to promote weight loss.

Solution

Randomly selected control and experimental groups of 1000 volunteers are studied over two years. Both receive pills: The experimental group receives the drug; the control group unknowingly receives a placebo. Technicians, who do not know which group each person is in, record the weight of the subjects every month. Experimenters, who never meet or see the volunteers, analyze the evidence gathered.

COMMUNICATION example 2

Act as a referee (peer-reviewer) to critique the following experimental design including, if necessary, suggestions for improvement.

Ten volunteers are provided with their horoscope for the test day. The volunteers orally respond in a group to the question: Does this horoscope describe your personality and life situation?

Solution

This experimental design is very inadequate because:

- ten volunteers is a very small sample size; the number needs to be at least 100 or more, randomly selected from thousands
- to control the horoscope variable, all subjects should be provided with the same horoscope
- to control subject interaction and influence, subjects should be isolated from one another
- to provide for accurate reporting of the subjects' responses, investigators should make audio or audiovisual recordings

The goal of technological problem solving is to solve practical problems by developing or revising a product or a process. The product or the process must fulfill its function, but it is not essential to understand why or how it works. Products are evaluated based on criteria such as simplicity, reliability, and cost. Technological processes are also evaluated by their efficiency. Technological products and processes may have both intended and unintended consequences. Therefore, it is important that various perspectives, such as ecological, economic, and political, are used in any assessment. For example, chlorofluorocarbons may be simple and inexpensive to make, and they may be useful for a particular function, but their effect on the ozone layer in the upper atmosphere must also be considered. Processes such as the chlorine bleaching of wood pulp may be efficient, but they may adversely affect an ecosystem. We often look to technological fixes as solutions for problems. Ecological, economic, political, legal, ethical, and/or social efforts by individuals or groups can often lead to more sustaining solutions than quick technological fixes.

Chemistry has always been closely associated with technology. Part of technology is the laboratory equipment, processes, and procedures used in both chemical and technological research and development. In modern chemistry, simple equipment and processes, such as beakers and filtration, are still used but chemistry also depends on sophisticated technology, such as computers, to store and manipulate the evidence collected.

C.1 Technological Problem-Solving Model

Technological problem solving is similar in some ways to scientific problem solving but its purpose differs. A characteristic of technological problem solving is a systematic, trial-and-error manipulation of variables (**Figure 3**). Variables are predicted and tested and the results are evaluated. When the cycle is repeated many times, the most effective set of conditions can be determined. Compare this model with the scientific problem-solving model in **Figure 1**, on page 790.

Technological problem-solving contexts include industrial, commercial, and consumer. The general process for technological problem solving is similar in all of these contexts. Technological problem solving is very common to us in our everyday lives. Learning more about this systematic trial-and-error approach can help us on an everyday basis.

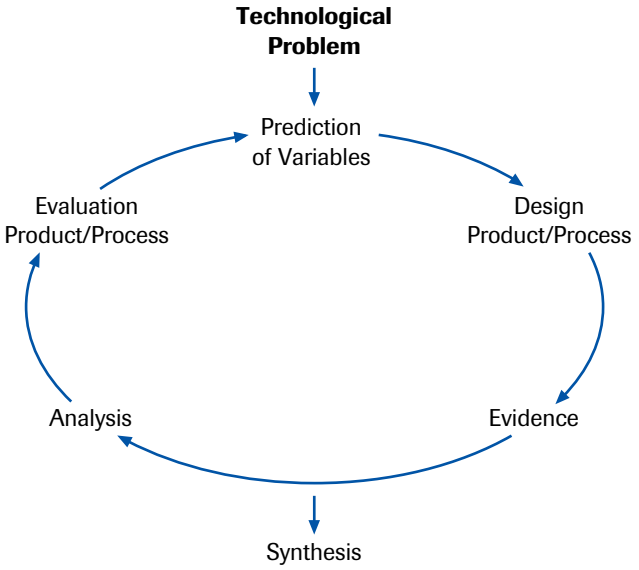


Figure 3
A technological problem-solving model

C.2 Investigation Reports

Investigation reports for technological problem-solving investigations can use the same headings as those for scientific problem-solving investigations (Appendix B.2, page 790). Some key differences between these two types of reports are

Purpose

The purpose will be to solve a specific, practical problem by developing or revising a product or process.

Evaluation

Evaluation criteria can be many and varied. In some cases, it will be sufficient to demonstrate that the product or process works for the materials used. You may also be asked to judge the simplicity, reliability, and efficiency of the product or process, recognize its value and limitations, and evaluate it from a variety of perspectives (Appendix D.2, page 806).

C.3 Laboratory Equipment

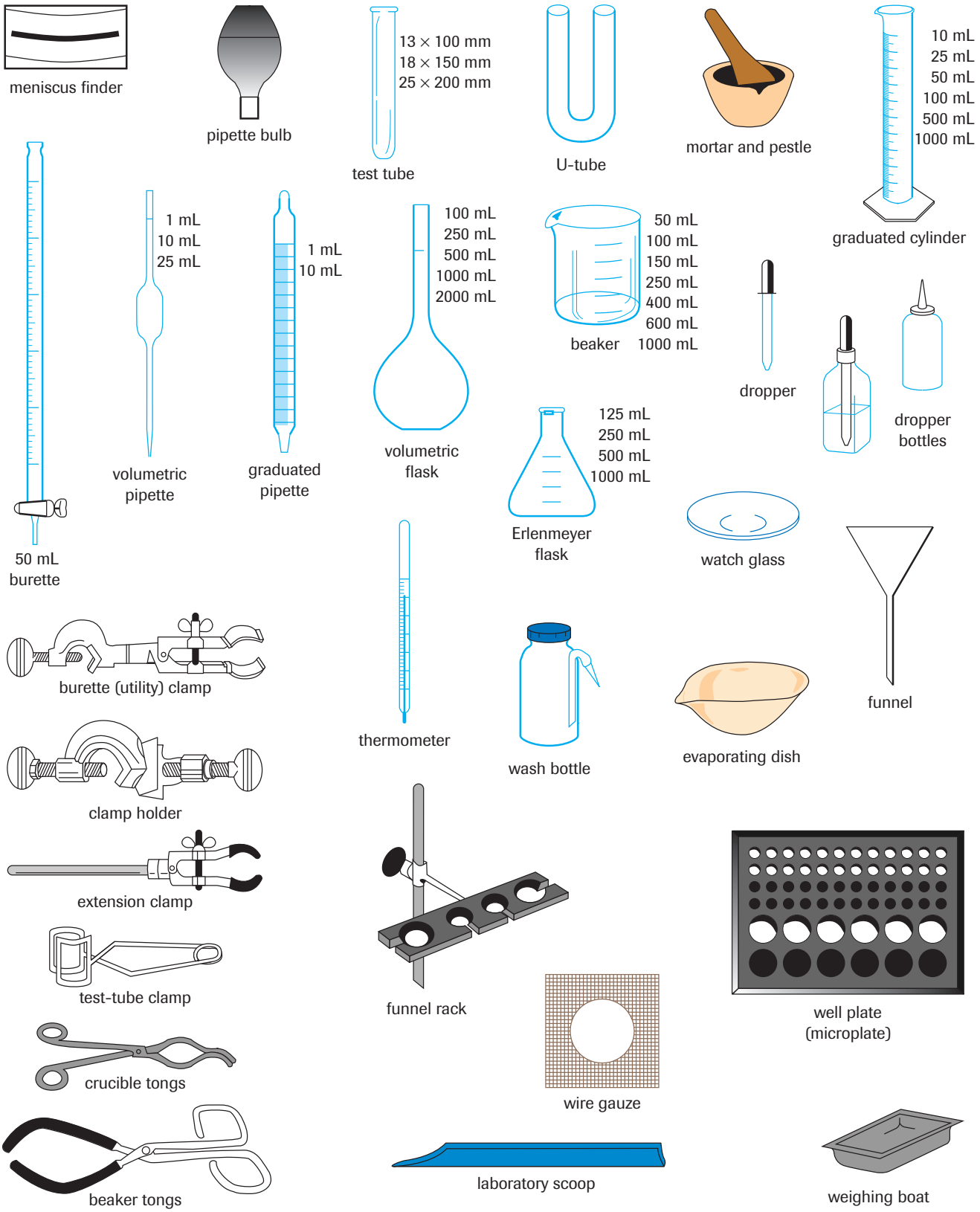


Figure 4
Common lab equipment

Glassware is breakable and should always be handled with care.

Using a Laboratory Burner

The procedure outlined below should be practised and memorized. Note the safety caution. You are responsible for your safety and the safety of others near you.

1. Turn the air and gas adjustments to the off position (**Figure 5**).
2. Connect the burner hose to the gas outlet on the bench.
3. Turn the bench gas valve to the fully on position.
4. If you suspect that there may be any gas leaks, replace the burner. (Give the leaky burner to your teacher.)

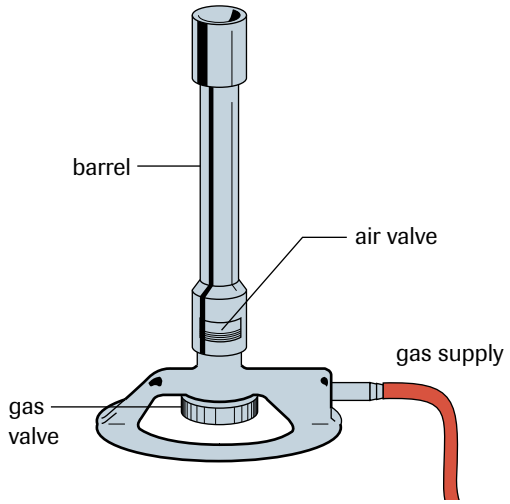


Figure 5
The parts of a common laboratory burner

5. While holding a lit match above and to one side of the barrel, open the burner gas valve until a small yellow flame results (**Figure 6**). If a striker is used instead of matches, generate sparks over the top of the barrel (**Figure 7**).
6. Adjust the air flow and obtain a pale blue flame with a dual cone (**Figure 8**). In most common types of laboratory burners, rotating the barrel adjusts the air intake. Rotate the barrel slowly. If too much air is added, the flame may go out. If this happens, immediately turn the gas flow off and relight the burner following the procedure outlined above. If your burners have a different kind of air adjustment, revise the procedure accordingly.
7. Adjust the gas valve on the burner to increase or decrease the height of the blue flame. The hottest part of the flame is the tip of the inner blue cone. Usually a 5 to 10 cm flame, which just about touches the object heated, is used.
8. Laboratory burners, when lit, should not be left unattended. If the burner is on but not being used, adjust the air and gas intakes to obtain a small yellow flame. This flame is more visible and, therefore, less likely to cause problems.

When lighting or using a laboratory burner, never position your head or fingers directly above the barrel. Tie back long hair and sleeves.

Figure 6
A yellow flame is a relatively cool flame and is easier to obtain than a blue flame when lighting a burner. A yellow flame is not used for heating objects because it contains a lot of black soot.

Figure 7
To generate a spark with a striker, pull up and across on the side of the handle containing the flint.

Figure 8
A pale, almost invisible flame is much hotter than a yellow flame. The hottest point is at the tip of the inner blue cone.

Using a Laboratory Balance

There are two types of balances: electronic (**Figure 9**) and mechanical (**Figure 10**). All balances must be handled carefully and kept clean. Always place chemicals into a container such as a beaker or plastic boat to avoid contamination and corrosion of the balance pan. To avoid error due to convection currents in the air, allow hot or cold samples to return to room temperature before placing them on the balance. Always record masses showing the correct precision. On a centigram balance, mass is measured to the nearest hundredth of a gram (0.01 g). When it is necessary to move a balance, hold the instrument by the base and steady the beam. Never lift a balance by the beams or pans.

Figure 9
An electronic balance

To avoid contaminating a whole bottle of reagent, do not scoop diectly from the original container of a chemical. Pour a quantity of the chemical into a clean, dry beaker or bottle, from which samples can be taken. Another acceptable technique for dispensing a small quantity of chemical is to rotate or tap the chemical bottle.

Using an Electronic Balance

Electronic balances are sensitive to small movements and changes in level; do not lean on the counter when using the balance.

1. Place a container or weighing paper on the balance.
2. Reset (tare) the balance so the mass of the container registers as zero.
3. Add chemical until the desired mass of chemical is displayed. Air currents or the high sensitivity of the balance may cause the last digit to vary.
4. Remove the container and sample.

There is a video demonstration of this technique on the Nelson Web site.

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Using a Mechanical Balance

Different kinds of mechanical balances are shown in **Figures 10(a)** and **(b)**. Some general procedures apply to most of them.

1. Clean and zero the balance. (Turn the zero adjustment screw so that the beam is balanced when the instrument reads 0 g and no load is on the pan.)
2. Place the container on the pan.
3. Move the largest beam mass one notch at a time until the beam drops, and then move the mass back one notch.
4. Repeat this process with the next smaller mass and continue until all masses have been moved and the beam is balanced. If you are using a dial type balance, the final step will be to turn the dial until the beam balances, as shown in **Figure 10(c)**.
5. Record the mass of the container.
6. Set the masses on the beams to correspond to the total mass of the container plus the desired sample.
7. Add the chemical until the beam is once again balanced.
8. Remove the sample from the pan and return balance to the zero position.

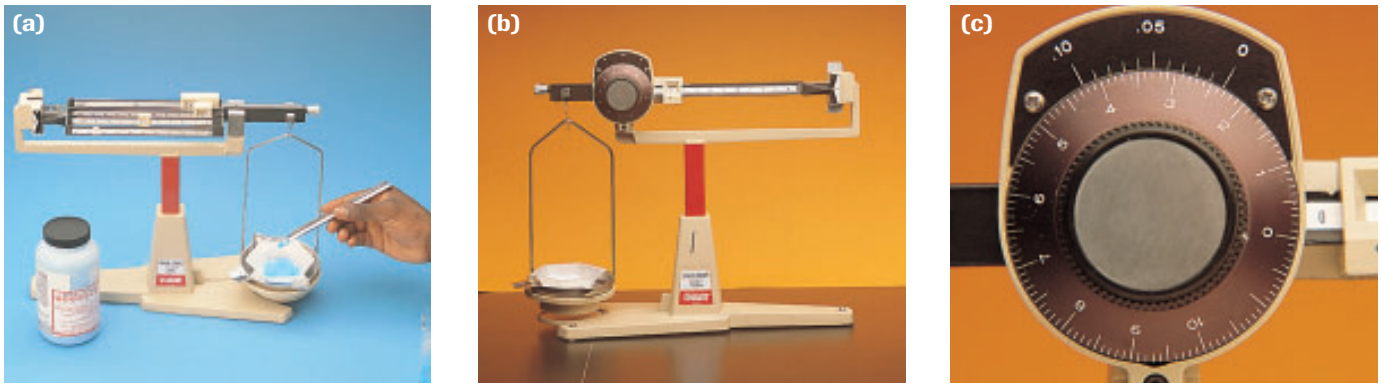

Figure 10
(a) On this type of mechanical balance, the sample is balanced by moving masses on several beams.
(b) Another type of mechanical balance has beams for the larger masses and a dial for the final adjustment.
(c) The dial reading on this balance with a vernier scale is 2.34 g. To read the hundredth of a gram, look below the zero on the vernier, and then look for the line on the vernier that lines up best with a line on the dial.

Using a Multimeter

A multimeter (**Figure 11**) is a device that measures a variety of electrical quantities, such as resistance, voltage, and current.

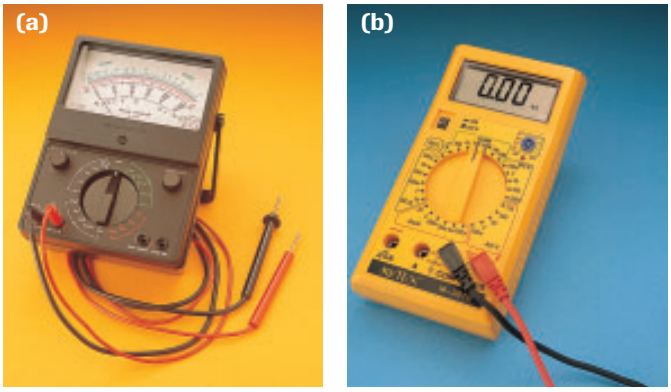

Figure 11
(a) An analog meter has a needle that moves in front of a labelled scale.
(b) A digital meter gives a direct reading with appropriate units.

Conductivity Measurements of Solutions

1. Set the dial on the meter to one of the higher values on the ohm (Ω) scale; for example, R × 100 or R × 1 K.
2. Touch the two metal probes together to check the battery. If the needle does not deflect significantly (more than one-half scale), have your teacher adjust the meter or replace the battery.
3. Test a sample of pure water as a control and note the movement of the needle.
4. Test your aqueous sample and record the deflection of the needle according to your teacher's instructions.
5. Rinse the probes with pure water before testing another sample.
6. Shut off the meter by using either the on/off switch or by turning the dial to any setting other than "Resistance."

Voltage Measurements of Cells and Batteries

1. Set the dial to the appropriate value on the direct current volts (DCV) scale; for example, 3 V.
2. The black lead (labelled negative or COM) is normally connected to the anode, and the red lead (positive) is connected to the cathode of a voltaic cell.
3. Make a firm contact between each metal probe and an electrode of the cell. (Press firmly with the pointed probe or use leads with an alligator clip.)
4. On analog meters (those with a needle), read the scale corresponding to the meter value you set in step 1.
5. If the needle attempts to move to the left off the scale or a digital meter registers a negative number, then switch the connections to the cell.

Using a Pipette

A pipette is a specially designed glass tube used to measure precise volumes of liquids. There are two types of pipettes and a variety of sizes for each type. A *volumetric pipette* (**Figure 12**) transfers a fixed volume, such as 10.00 mL or 25.00 mL, accurate to within 0.04 mL. A *graduated pipette* (**Figure 13**) measures a range of volumes, just as a graduated cylinder does. A 10 mL graduated pipette delivers volumes accurate to within 0.1 mL. There is a video demonstration of this technique on the Nelson Web site.

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Figure 12
A volumetric pipette delivers the volume printed on the label if the temperature is near room temperature.

Figure 13
To use a graduated pipette, you must be able to start and stop the flow of the liquid.

1. Rinse the pipette with small volumes of distilled water using a wash bottle, and then with the sample solution.
   A clean pipette has no visible residue or liquid drops clinging to the inside wall. Rinsing with aqueous ammonia and scrubbing with a pipe cleaner might be necessary to clean the pipette.
2. Hold the pipette with your thumb and fingers near the top. Leave your index finger free.
3. Place the pipette in the sample solution, resting the tip on the bottom of the container if possible. Be careful that the tip does not hit the sides of the container.
4. Squeeze the bulb into the palm of your hand and place the bulb firmly and squarely on the end of the pipette (**Figure 14**) with your thumb across the top of the bulb.
5. Release your grip on the bulb until the liquid has risen above the calibration line.
   This may require bringing the level up in stages: remove the bulb, put your finger on the pipette, squeeze the air out of the bulb, re-place the bulb, and continue the procedure.
6. Remove the bulb, placing your index finger over the top.
   If you are using a dispensing bulb (**Figure 15**), it remains attached to the pipette.
7. Wipe all solution from the outside of the pipette using a paper towel.
8. While touching the tip of the pipette to the inside of a waste beaker, gently roll your index finger (or squeeze the valve of the dispensing bulb) to allow the liquid level to drop until the bottom of the meniscus reaches the calibration line (**Figure 16**).
   To avoid parallax errors, set the meniscus at eye level. Stop the flow when the bottom of the meniscus is on the calibration line. Use the bulb to raise the level of the liquid again if necessary.
9. While holding the pipette vertically, touch the pipette tip to the inside wall of a clean receiving container. Remove your finger or adjust the valve and allow the liquid to drain freely until the solution stops flowing.
10. Finish by touching the pipette tip to the inside of the container held at about a 45° angle (**Figure 17**). Do not shake the pipette. The delivery pipette is calibrated to leave a small volume in the tip.

Never use your mouth to draw a liquid up a pipette. Always use a pipette bulb.

Figure 14
Release the bulb slowly. Pressing down with your thumb placed across the top of the bulb maintains a good seal. Setting the pipette tip on the bottom slows the rise or fall of the liquid.

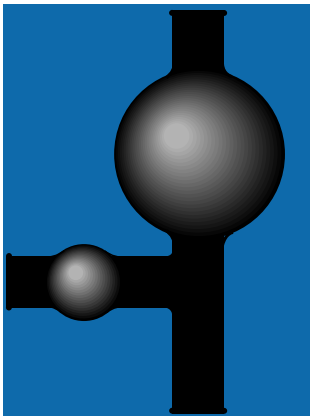

Figure 15
A dispensing pipette bulb uses a small valve in the side stem to control the flow of liquid in a pipette.

Figure 16
To allow the liquid to drop slowly to the calibration line, it is necessary for your finger and the pipette top to be dry. Also keep the tip on the bottom to slow down the flow.

Figure 17
A vertical volumetric pipette is drained by gravity and then the tip is placed against the inside wall of the container. A small volume is expected to remain in the tip.

C.4 Laboratory Processes

The processes or experimental procedures listed below are part of common designs used in scientific or technological laboratories.

Crystallization

Crystallization is used to separate a solid from a solution by evaporating the solvent or lowering the temperature. Evaporating the solvent is useful for quantitative analysis of a binary solution; lowering the temperature is commonly used to purify and separate a solid whose solubility is temperature-sensitive. Chemicals that have a low boiling point or decompose on heating cannot be separated by crystallization using a heat source.

1. Measure the mass of a clean beaker or evaporating dish.
2. Place an accurate volume of the solution in the container.
3. Set the container aside to evaporate the solution slowly, or warm the container gently on a hot plate or with a laboratory burner.
4. When the contents appear dry, measure the mass of the container and solid (**Figure 18**).

Figure 18
When the substance has crystallized, it may appear dry but small quantities of water may still be present. To be certain the solid is dry, it must be heated until the mass becomes constant.

5. Heat the solid with a hot plate or burner, cool it, and measure the mass again.
6. Repeat step 5 until the final mass remains constant. (Constant mass indicates that all of the solvent has evaporated.)

Filtration

In filtration, solid is separated from a mixture using a porous filter paper. The more porous papers are called qualitative filter papers. Quantitative filter papers allow only invisibly small particles through the pores of the paper. There is a video demonstration of this technique on the Nelson Web site.

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1. Set up a filtration apparatus (**Figure 19**): stand, funnel holder, filter funnel, waste beaker, wash bottle, and a stirring rod with a flat end for scraping.

Figure 19
The tip of the funnel should touch the inside wall of the collecting beaker.

2. Fold the filter paper along its diameter, and then fold it again to form a cone. A better seal of the filter paper on the funnel is obtained if a small piece of the outside corner of the filter paper is torn off (**Figure 20**).

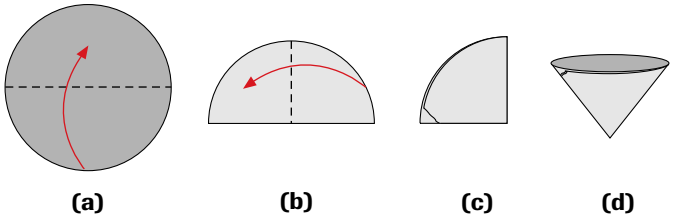

Figure 20
To prepare a filter paper, fold it in half twice, and then remove the outside corner as shown.

3. Measure and record the mass of the filter paper after removing the corner.
4. While holding the open filter paper in the funnel, wet the entire paper and seal the top edge firmly against the funnel.

5. With the stirring rod touching the spout of the beaker, decant most of the solution into the funnel (**Figure 21**). Transferring the solid too soon clogs the pores of the filter paper. Keep the level of liquid about two-thirds up the height of the filter paper. The stirring rod should be rinsed each time it is removed.

Figure 21
Pouring along the stirring rod prevents drops of liquid from going down the outside of the beaker when you stop pouring.

6. When most of the solution has been filtered, pour the remaining solid and solution into the funnel. Use the wash bottle and the flat end of the stirring rod to clean any remaining solid from the beaker.
7. Rinse the stirring rod and the beaker.
8. Wash the solid two or three times to ensure that no solution is left in the filter paper. Direct a gentle stream of water around the top of the filter paper.
9. When the filtrate has stopped dripping from the funnel, remove the filter paper. Press your thumb against the thick (three-fold) side of the filter paper and slide the paper up the inside of the funnel.
10. Transfer the filter paper from the funnel onto a labelled watch glass and unfold the paper to let the precipitate dry.
11. Determine the mass of the filter paper and dry precipitate.

Preparation of Standard Solutions

Laboratory procedures often call for the use of a solution of specific, accurate concentration. The apparatus used to prepare such a solution is a volumetric flask. A meniscus finder is useful in setting the bottom of the meniscus on the calibration line (**Figure 22**).

Preparing a Standard Solution from a Solid Reagent

There is a video demonstration of this technique on the Nelson Web site.

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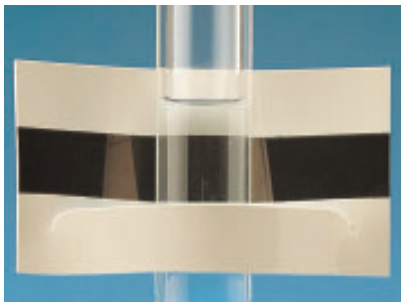

Figure 22
Raise the meniscus finder along the back of the neck of the volumetric flask until the meniscus is outlined as a sharp, black line against a white background.

1. (Prelab) Calculate the required mass of solute from the volume and concentration of the solution.
2. Measure the required mass of solute in a clean, dry beaker or weighing boat. (Refer to "Using a Laboratory Balance" on page 799.)
3. Pour less than one-half of the final volume of pure water into a beaker. Transfer the solute to the water. Stir to dissolve.
4. Transfer the solution and all water used to rinse the equipment into a clean volumetric flask. (The beaker and any other equipment should be rinsed two or three times with pure water.)
5. Add pure water, using a medicine dropper for the final few millilitres while using a meniscus finder to set the bottom of the meniscus on the calibration line.
6. Stopper the flask and mix the solution by slowly inverting the flask several times.

Preparing a Standard Solution by Dilution

There is a video demonstration of this technique on the Nelson Web site.

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1. (Prelab) Calculate the volume of concentrated reagent required.
2. Add approximately one-half of the final volume of pure water to the volumetric flask.
3. Measure the required volume of stock solution using a pipette. (Refer to "Using a Pipette" on page 800).
4. Transfer the stock solution slowly into the volumetric flask while mixing.
5. Add pure water, and then use a medicine dropper and a meniscus finder to set the bottom of the meniscus on the calibration line (**Figure 22**).

6. Stopper and mix the solution by slowly inverting the flask several times.

If water is added directly to some solids or concentrated liquids, there may be boiling or splattering. Always add a solid solute or concentrated liquids to water.

Titration

Titration is used in the volumetric analysis of an unknown concentration of a solution. Titration involves adding a solution (the titrant) from a burette to another solution (the sample) in an Erlenmeyer flask until a recognizable endpoint, such as a colour change, occurs. (See the video on the Nelson Web site.)

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1. Rinse the burette with small volumes of pure water using a wash bottle. Using a burette funnel, rinse with small volumes of the titrant (**Figure 23**). (If liquid droplets remain on the sides of the burette after rinsing, scrub the burette with a burette brush. If the tip of the burette is chipped or broken, replace the tip or the whole burette.)

Figure 23
A burette should be rinsed with water and then the titrant before use.

2. Using a small burette funnel, pour the titrant solution into the burette until the level is near the top. Open the stopcock for maximum flow to clear any air bubbles from the tip and to bring the liquid level down to the scale.
3. Record the initial burette reading to the nearest 0.1 mL. Avoid parallax errors by reading volumes at eye level with the aid of a meniscus finder.
4. Pipette a known volume of the solution of unknown concentration into a clean Erlenmeyer flask. Place a white piece of paper beneath the Erlenmeyer flask to make it easier to detect colour changes.
5. Add an indicator if one is required. Add the smallest quantity necessary (usually 1 to 2 drops) to produce a noticeable colour change in your sample.
6. Add the solution from the burette quickly at first, and then slowly, drop-by-drop, near the endpoint (**Figure 24**). Stop as soon as a drop of the titrant produces a permanent colour change in the sample solution. A permanent colour change is considered to be a noticeable change that lasts for 10 s after swirling.
7. Record the final burette reading to the nearest 0.1 mL.
8. The final burette reading for one trial becomes the initial burette reading for the next trial. Three trials with results within 0.2 mL are normally required for a reliable analysis of an unknown solution.
9. Drain and rinse the burette with pure water. Store the burette upside down with the stopcock open.

Figure 24
Near the endpoint, continuous gentle swirling of the solution is particularly important.

Diagnostic Tests

The tests described in **Table 1** are commonly used to detect the presence of a specific substance. All diagnostic tests include a brief procedure, some expected evidence, and an interpretation of the evidence obtained. This is conveniently communicated using the format, "If [procedure] and [evidence], then [analysis]." Diagnostic tests can be constructed using any characteristic empirical property of a substance. For example, diagnostic tests for acids, bases, and neutral substances can be specified in terms of the pH of the solutions. For specific chemical reactions, properties of the products that the reactants do not have, such as the insolubility of a precipitate, the production of a gas, or the colour of ions in aqueous solutions, can be used to construct diagnostic tests.

If possible, you should use a control to illustrate that the test does not give the same results with other substances. For example, in the test for oxygen, inserting a glowing splint into a test tube that contains only air is used to compare the effect of air on the splint with a test tube in which you expect oxygen has been collected.

Communication of Diagnostic Tests

The procedure, evidence, and analysis information for a diagnostic test can be communicated in three different formats:

- "If ... and ... then ..." statement
- table
- flowchart

Table 1 Some Standard Diagnostic Tests

Substance tested	Diagnostic test
water	If cobalt(II) chloride paper is exposed to a liquid or vapour, and the paper turns from blue to pink, then water is likely present.
oxygen	If a glowing splint is inserted into the test tube, and the splint glows brighter or re-lights, then oxygen gas is likely present.
hydrogen	If a flame is inserted into the test tube, and a squeal or pop is heard, then hydrogen is likely present.
carbon dioxide	If the unknown gas is bubbled into a limewater solution, and the limewater turns cloudy, then carbon dioxide is likely present.
halogens	If a few millilitres of a hydrocarbon solvent is added, with shaking, to a solution in a test tube, and the colour of the solvent appears to be • light yellow-green, then chlorine is likely present • orange, then bromine is likely present • purple, then iodine is likely present
acid	If strips of blue and red litmus paper are dipped into the solution, and the blue litmus turns red, then an acid is present.
base	If strips of blue and red litmus paper are dipped into the solution, and the red litmus turns blue, then a base is present.
neutral solution	If strips of blue and red litmus paper are dipped into the solution, and neither litmus changes colour, then only neutral substances are likely present.
neutral ionic solution	If a neutral solution is tested for conductivity with a multimeter, and the solution conducts a current, then a neutral ionic substance is likely present.
neutral molecular solution	If a neutral solution is tested for conductivity with a multimeter, and the solution does not conduct a current, then a neutral molecular substance is likely present.
	There are thousands of diagnostic tests. You can create some of these using data from the periodic table (on the inside front cover of this book), and from the data tables in Appendix I and on the inside back cover.

Science is a human endeavour, technology has a social purpose, and both have always been part of society. Science and technology together affect society in a myriad of ways. Society also affects science and technology by placing controls on them and expecting solutions to societal problems.

When controversial issues related to science and technology arise in our society, there is often heated debate among various special-interest groups. Often, little progress is made because different parties in the debate generally recognize only a single perspective on the issue. Many people now realize that an informed multi-perspective view is more defensible. The following model represents one possible procedure for making an informed decision on a social issue related to science and technology.

D.1 STS Decision-Making Model

1. ***Identify an STS (science–technology–society) issue.*** Newspapers, magazines, and news broadcasts are sources of current STS issues. However, some issues like acid rain have been current for some time and rarely appear in the news. When identifying an issue for debate, it is convenient to state the issue as a resolution (e.g., "Be it resolved that the use of fossil fuels for heating homes should be eliminated.").
2. ***Design a plan to address the STS issue.*** Possible designs include individual research, a debate, a town-hall meeting (or role-playing), or participation in an actual hearing or on a committee.
3. ***Identify and obtain relevant information on as many perspectives as possible.*** An STS issue will always have scientific and technological perspectives. Common perspectives are shown in **Table 2**.

Table 2 Perspectives on STS Issues

scientific	ethical
technological	social
ecological	militaristic
economic	esthetic
political	mystical
legal	emotional

Another perspective is the world view or perspective of Aboriginal peoples. In general, Aboriginal peoples believe that we are an integral part of our environment. Their holistic view includes not only a physical interdependence but also a spiritual one.

Information from different perspectives can be obtained from references and through group discussions. There are many sides to every issue. There can be positive and negative viewpoints about the resolution from every perspective.

4. ***Generate a number of alternative solutions to the STS problem.*** Some obvious solutions will arise from the resolution. Other creative solutions often arise from a brainstorming session within a group.
5. ***Evaluate each solution and decide which is best.*** One method is to rank the value of a particular solution from each perspective. For example, a solution might have little economic advantage and be ranked as 1 on a scale of 1 to 5; the solution might have a significant ecological benefit and be ranked as 5, for a total of 6. A different solution might be judged as 3 from the economic perspective and 1 from the ecological perspective, for a total of 4. The solution with the highest total is likely to be chosen. Although simplistic, this method facilitates evaluation and illustrates the tradeoffs that occur in any real issue.

D.2 Types of Reports

There are many ways to communicate the results of an investigation of an STS issue (**Table 3**). All methods will require some research about the issue and perspectives on the issue (including positive and negative viewpoints). Some methods can also include alternative solutions and the evaluation of these solutions. Working within a group and brainstorming is a useful process. No matter how the issue will be presented and reported, you need to be well prepared.

Table 3 STS Investigations and Reports

Plan	Reporting suggestions
individual or group research	• written report or poster • multimedia presentation
debate	• research notes • videotape of the debate
role-playing (e.g., town hall meeting)	• research notes • videotape of the meeting
survey	• survey form with tables and graphs
newspaper article	• published article

E.1 Laboratory Safety

Safety is always important in a laboratory or in other settings that feature chemicals or technological devices. It is your responsibility to be aware of possible hazards, to know the rules—including ones specific to your classroom—and to behave appropriately. Always alert the teacher in case of any accident.

Alberta Education has an extensive document, "Safety in the Science Classroom," that deals with hazards and safety.

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General Safety Rules

Safety in the laboratory is an attitude and a habit more than it is a set of rules. It is easier to prevent accidents than to deal with the consequences of an accident. Most of the following rules are common sense:

- Read all directions before doing any laboratory work, and follow all verbal instructions.
- Know the potential hazards, including the contents and location of MSDS, and the location of all safety equipment.
- Wear eye protection and lab aprons/coats.
- Behave responsibly. Avoid sudden or rapid motion that may interfere with someone carrying or working with chemicals.
- Wear closed shoes (not sandals or bare feet) when working in the laboratory.
- Place your books, bags, and purses away from the work area.
- Do not chew gum, eat, drink, or taste anything in the laboratory.
- Ask for assistance when you are not sure how to do a procedural step.
- Inform your teacher immediately if any problem or accident occurs.
- Never attempt any unauthorized or unsupervised experiments.
- Never handle any chemical with your hands. Use a laboratory scoop or spoon for solids.
- Never use the contents of a bottle that has no label or an illegible label. Always double check the label to ensure that you are using the chemical needed. Always pour from the side opposite the label.
- When leaving chemicals in containers, ensure that they are labelled.
- Do not take any more chemical than needed and never return excess chemicals to their original container.
- Hold larger bottles with both hands; one hand on the base.
- Do not inhale any vapours directly from any container. If smell is to be tested, fan the vapours toward your nose, keeping the container away from underneath your nose.
- Always use a pipette bulb, and never pipette by mouth.
- When heating a test tube over a burner, use a test-tube holder with the test tube at an angle, facing away from you and others. Gently move the test tube backwards and forwards through the flame.
- Clean up all spills, even spills of water, immediately. Clean up your work area at the end of an experiment.
- Dispose of chemicals appropriately as directed by your teacher.
- Always wash your hands with soap and water before you leave the laboratory.
- Do not forget safety procedures when you leave the laboratory. These same rules also apply at home or at work.

Glass Safety and Cuts

- Never use glassware that is cracked or chipped. Give such glassware to your teacher or dispose of it as directed. Do not put the item back into circulation.
- Never pick up broken glassware with your fingers. Use a broom and dustpan.
- Do not put broken glassware into garbage containers. Dispose of glass fragments in special containers marked "broken glass."
- If you cut yourself, inform your teacher immediately. Imbedded glass or continued bleeding requires medical attention.

Burns

- In a laboratory where burners or hot plates are being used, never pick up a glass object without first checking the temperature by lightly and quickly touching the item. Glass items that have been heated stay hot for a long time but do not appear to be hot. Metal items such as ring stands and hot plates can also cause burns; take care when touching them.
- Before using a laboratory burner, make sure that long hair is always tied back. Do not wear loose clothing. (Wide long sleeves should be tied back or rolled up.)

- Do not use a laboratory burner near wooden shelves, flammable liquids, or any other item that is combustible. Know how to use the type of burner in your laboratory. (See Using a Laboratory Burner, page 798)
- Never look down the barrel of a laboratory burner.
- Always pick up a burner by the base, never by the barrel.
- Never leave a lighted bunsen burner unattended.
- If you burn yourself, *immediately* run cold water over the burned area and inform your teacher.

Eye Safety

- Always wear approved eye protection in a laboratory, no matter how simple or safe the task appears to be. Keep the safety glasses over your eyes, not on top of your head. For certain experiments, full face protection may be necessary.
- Never look directly into the opening of flasks or test tubes.
- If, in spite of all precautions, you get a solution in your eye, quickly use the eyewash station or nearest running water. Continue to rinse the eye with water for at least 15 minutes. This is a very long time—have someone time you. Unless you have a plumbed eyewash system, you will also need assistance in refilling the eyewash container. Have another student inform your teacher of the accident. The injured eye should be examined by a doctor.
- It is recommended that you do not wear contact lenses in the laboratory. If you wear contact lenses in the laboratory, there is a danger that a chemical might get behind the lens where it cannot be rinsed out with water. If you must wear contact lenses in the chemistry laboratory, be extra careful. Tell your teacher if you are wearing contact lenses in the laboratory. Whether or not you wear contact lenses, do not touch your eyes without first washing your hands.
- If a piece of glass or other foreign object enters an eye, immediate medical attention is required.

Fire Safety

Immediately inform your teacher of any fires. Very small fires in a container may be extinguished by covering the container with a wet paper towel or a ceramic square, which would cut off the supply of air. If anyone's clothes or hair catch fire, the fire can be extinguished by smothering the flames with a blanket or a piece of clothing. Larger fires require a fire extinguisher. (Know how to use the fire extinguisher that is in your laboratory.) If the fire is too large to approach safely with an extinguisher, vacate the location and sound the fire alarm. (School staff will inform the fire department.)

If you use a fire extinguisher, direct the extinguisher at the base of the fire and use a sweeping motion, moving the extinguisher nozzle back and forth across the front of the fire's base. You must use the correct extinguisher for the kind of fire you are trying to control. Each extinguisher is marked with the class of fire for which it is effective. The fire classes are outlined below. Most fire extinguishers in schools are of the ABC type.

- Class A fires involve ordinary combustible materials that leave coals or ashes, such as wood, paper, or cloth. Use water or dry chemical extinguishers on Class A fires. (Carbon dioxide extinguishers are not satisfactory as carbon dioxide dissipates quickly and the hot coals can reignite.)
- Class B fires involve flammable liquids such as gasoline or solvents. Carbon dioxide or dry chemical extinguishers are effective on Class B fires. (Water is not effective on a Class B fire since the water splashes the burning liquid and spreads the fire.)
- Class C fires involve live electrical equipment, such as appliances, photocopiers, computers, or laboratory electrical apparatus. Carbon dioxide or dry chemical extinguishers are recommended for Class C fires. Carbon dioxide extinguishers are much cleaner than the dry chemical variety. (Using water on live electrical devices can result in severe electrical shock.)
- Class D fires involve burning metals, such as sodium, potassium, magnesium, or aluminium. Sand or salt are usually used to put out Class D fires. (Using water on a metal fire can cause a violent reaction.)
- Class E fires involve a radioactive substance. These involve special considerations at each site.

Electrical Safety

Water or wet hands should never be used near electrical equipment. When unplugging equipment, remove the plug gently from the socket (do not pull on the cord). Do not use any devices with electric motors when flammable liquids are present unless the area is well ventilated.

E.2 Safety Symbols and Information

Educational, Commercial, and Industrial Information

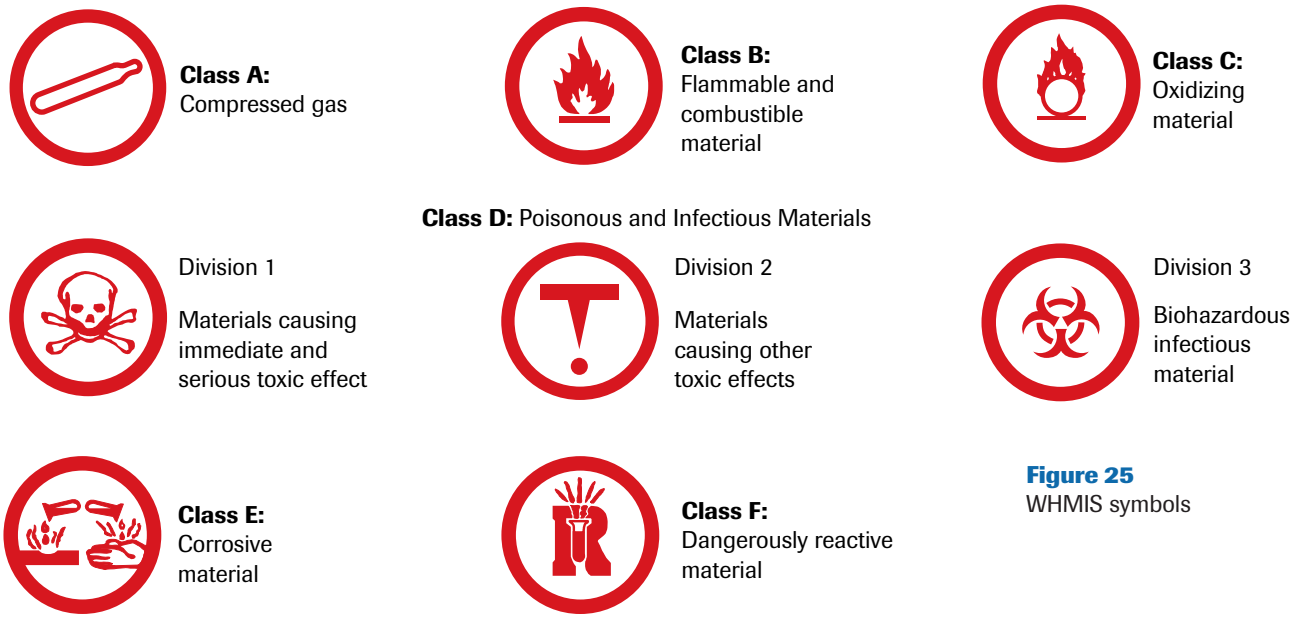
Class A: Compressed gas
Class B: Flammable and combustible material
Class C: Oxidizing material
Class D: Poisonous and Infectious Materials
Division 1
Materials causing immediate and serious toxic effect
Division 2
Materials causing other toxic effects
Division 3
Biohazardous infectious material
Class E: Corrosive material
Class F: Dangerously reactive material

Figure 25
WHMIS symbols

The Workplace Hazardous Materials Information System (WHMIS) provides workers and students with complete and accurate information regarding hazardous products. All chemical products supplied to schools, businesses, and industry must contain standardized labels and be accompanied by Material Safety Data Sheets (MSDS) providing detailed information about the product. Clear and standardized labelling is an important component of WHMIS (**Figure 25**). These labels must be present on the product's original container or be added to other containers if the product is transferred.

Although MSDS must be supplied with every product sold, current MSDS can also be obtained at several Internet sites, which are useful for researching information about chemicals.

www.science.nelson.com GO

Consumer Information

The Canadian Hazardous Products Act requires manufacturers of consumer products containing chemicals to include a symbol specifying both the nature and degree of the primary hazard, and to note any secondary hazards, first aid treatment, storage, and disposal. The symbols show the hazard by an illustration and the degree of the hazard by the type of border surrounding the illustration (**Figure 26**).

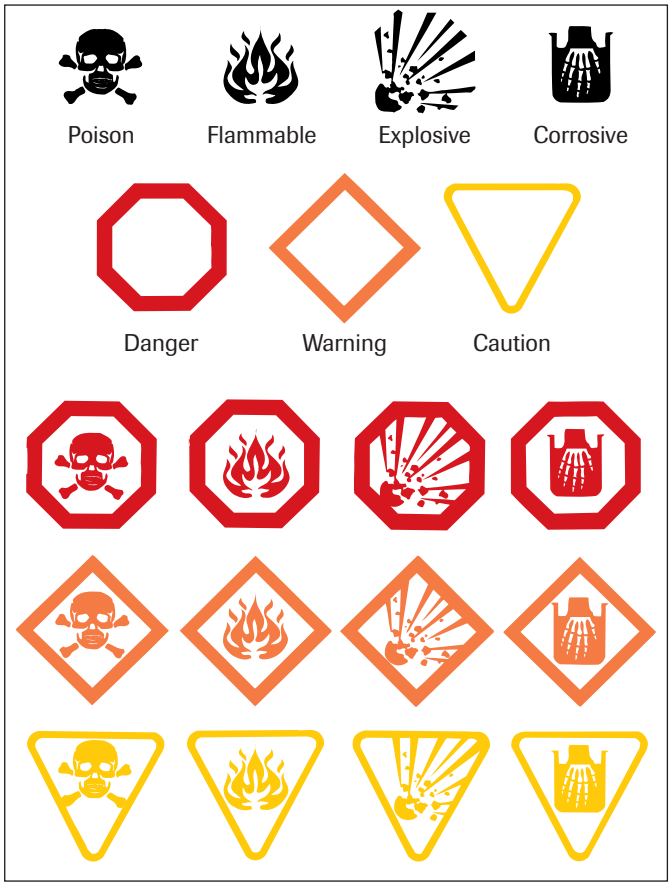

Figure 26
Household Hazardous Product Symbols

E.3 Waste Disposal

Disposal of chemical wastes at home, at school, or at work is a societal issue. We all need to be stewards of our planet; in other words, to behave as custodians or keepers. Some governments, institutions, and industries have begun to implement product stewardship programs. This is an environmental management plan based on the principle that whoever designs, produces, sells, or uses a product should take responsibility for minimizing the product's environmental impact over its complete life cycle. Governments have regulations for the handling, transportation, and disposal of chemicals, but each of us needs to take responsibility for the wastes we produce at home and at school.

Most laboratory waste can be washed down the drain, or, if it is in solid form, placed in ordinary garbage containers. However, some waste must be treated more carefully. Throughout this textbook, special waste disposal problems are noted, but it is your responsibility to dispose of waste in the safest possible manner.

Flammable Substances

Flammable liquids should not generally be washed down the drain. (The exceptions to this rule are aqueous solutions of non-toxic flammables such as alcohol–water solutions: they can safely be flushed.) Special fire-resistant containers are used to store flammable liquid waste. Waste solids that pose a fire hazard should be stored in fireproof containers. Care must be taken not to allow flammable waste to come into contact with any sparks, flames, other ignition sources, or oxidizing materials. The particular method of disposal depends on the nature of the substance.

Corrosive Solutions

Solutions that are corrosive but not toxic, such as acids, bases, or oxidizing agents, can usually be washed down the drain, but care should be taken to ensure that they are properly neutralized and diluted.

To neutralize diluted waste acids, use diluted waste bases, and vice versa. Or, use sodium bicarbonate for neutralizing the acid and use dilute hydrochloric acid for neutralizing the base. Oxidizing agents, such as potassium permanganate, should also be diluted with a 10% aqueous solution of sodium thiosulfate (reducing agent) before washing them into the drain.

Use large quantities of water and continue to pour water down the drain for a few minutes after all the substance has been washed away.

Heavy Metal Solutions

Heavy metal compounds (for example, lead, mercury, or cadmium compounds) should not be flushed down the drain. These substances are cumulative poisons and should be kept out of the environment. A special container is kept in the laboratory for heavy metal solutions. Pour any heavy metal waste into this container. Remember that paper towels used to wipe up solutions of heavy metals, as well as filter papers with heavy metal compounds imbedded in them, should be treated as solid toxic waste.

Disposal of heavy metal solutions is usually accomplished by precipitating the metal ion (for example, as lead(II) silicate) and disposing of the solid. Disposal may be by elaborate means such as deep well burial, or by simpler but accepted means such as delivering the substance to a landfill. Heavy metal compounds should not be placed in school garbage containers. Usually, waste disposal companies collect materials that require special disposal and dispose of them as required by law.

Toxic Substances

Solutions of toxic substances, such as oxalic acid, should not be poured down the drain, but should be disposed of in the same manner as heavy metal solutions. Solid toxic substances are handled similarly to precipitates of heavy metal. Chemicals should be stored in their original containers, with their labels clearly visible.

Communication is essential in science. The international scope of science requires that quantities, chemical symbols, and mathematical tools such as numbers, operations, tables, and graphs, be understood by scientists in different countries with different languages. The way in which scientific knowledge is expressed also reflects the nature of scientific knowledge, and in particular, the certainty of the knowledge.

F.1 Scientific Language

Science deals with two types of knowledge: empirical (observable "facts") and theoretical (non-observable ideas). Directly observable knowledge is generally considered to be more certain than interpretations or theoretical concepts. Theories are subject to change and, therefore, are less certain than the observations upon which they are based.

When observations are interpreted or explained, the language used should reflect some uncertainty or tentativeness. Use phrases such as

- The evidence suggests that…
- According to the theory of…
- It appears likely that…
- Scientists generally believe that…
- One could hypothesize that…

Avoid the use of the word "prove." Scientific concepts cannot be proven. The evidence may be extensive and reliable, but a concept to explain the evidence will never be 100% certain. In general, the language that you use should reflect the certainty of the information (observations are more certain than scientific concepts), and it should refer to the evidence available to you.

DID YOU KNOW ?

Confidence in Empirical versus Theoretical Knowledge

A candle does not burn unless air is present. In a closed container, a candle flame is extinguished after a short period of time. These are simple and relatively certain facts that can be directly stated. At one time, scientists believed that burning releases a substance called phlogiston, which was absorbed by the air until it could hold no more phlogiston; this is what stopped the burning. This theory, which was firmly believed by many chemists until the 1800s, was eventually replaced by the oxygen theory of combustion. The facts (evidence) remained the same but the idea (theory) completely changed.

F

F.2 SI Symbols and Conventions

The International System of Units, known as SI from the French name, *Système international d'unités*, is the measurement and communication system used internationally by scientists; it is also the legal measurement system in Canada and most countries in the world. Physical quantities are ultimately expressed in terms of seven fundamental SI units, called base units, which cannot be expressed as combinations of simpler units (**Table 4**). Although the base unit for mass is the kilogram (kg), it is more common in a chemistry laboratory to use the gram (g). Similarly, although the base unit for temperature (T) is kelvin (K), the common temperature (t) unit is degree Celsius (°C).

All other quantities can be expressed in terms of these seven fundamental quantities. For convenience, a unit derived from a combination of base units may be assigned a symbol of its own. **Table 5** lists a few of the physical quantities and derived units most commonly encountered in chemistry.

Table 4 Quantities and Fundamental Base Units

Quantity	Symbol	Unit	Symbol
length	l	metre	m
time	t	second	s
mass	m	kilogram	kg
chemical amount	n	mole	mol
temperature	T	kelvin	K
electric current	I	ampere	A
luminous intensity	I_v	candela	cd

Quantities and their SI base units are listed in Table 5, on the next page. These are the units most widely used by scientists. For convenience, however, units such as tonne (T) for *mass* and annum (a) for *year* are sometimes used to represent quantities that would be inconveniently large when expressed as base units.

Table 5 Quantities and Base Units

Quantity	Symbol	Unit	Symbol
molar mass	M	grams per mole	g/mol
volume	V	litre	L
amount concentration	c	moles per litre	mol/L
pressure	P	pascal	Pa
energy	E	joule	J
heat capacity	C	joules per degree Celsius	J/°C
specific heat capacity	c	joules per gram per degree Celsius	J/(g•°C)
volumetric heat capacity	c	megajoules per cubic metre per degree Celsius	MJ/(m^3•°C)
molar enthalpy	$\Delta_r H_m$	kilojoules per mole	kJ/mol
enthalpy change	$\Delta_r H$	kilojoules	kJ
electric charge	Q	coulomb	C
electric potential difference (voltage)	E	volt (joules per coulomb)	V

SI Prefixes

Next to universality, the most important feature of any system of units is convenience. SI has been designed to maximize convenience in a number of ways. A given quantity is always measured in the same base unit regardless of the context in which it is measured. For example, all forms of energy, including energy in food, are measured in joules. When a unit is too large or too small for convenient measurement, the unit is adjusted in size with a prefix. (See **Table 6**.) Prefixes allow units to be changed in size by multiples of ten. However, except for the use of "centi" in centimetre, we commonly use only prefixes that change the unit in multiples of a thousand.

Table 6 Some SI Prefixes

Prefix	Symbol	Fact
tera	T	10^{12}
giga	G	10^{9}
mega	M	10^{6}
kilo	k	10^{3}
milli	m	10^{-3}
micro	μ	10^{-6}
nano	n	10^{-9}
pico	p	10^{-12}

Scientific Notation

Scientific notation is a convenient method for expressing either a very large value or a very small value as a number between 1 and 10 multiplied by a power of 10. For example, the following numbers are expressed in regular notation and scientific notation:

Regular notation	Scientific notation
1200 L	1.200×10^3 L
0.000 000 998 mol/L	9.98×10^{-7} mol/L

On some calculators, the F ⇌ E key or the FSE key changes the number in the display into or from scientific notation. To enter a value in scientific notation in your calculator, the EXP or EE key is used to enter the power of ten. Note that the base 10 is not keyed into the calculator. For example, to enter

1.200×10^3 press [1] [•] [2] [EXP] [3]

9.98×10^{-7} press [9] [•] [9] [8] [EXP] [7] [+/–]

All mathematical operations and functions (such as +, –, ×, ÷, log) can be carried out with numbers in scientific notation.

Scientific notation is useful in calculations because it simplifies the cancellation of units and the totalling of powers of ten. However, scientific notation is sometimes overused. SI recommends that, wherever possible, prefixes be used to report measured values. Scientific notation should be reserved for situations where no prefix exists, or where it is essential to use the same unit (for example, comparing a wide range of energy values in kilojoules per gram). A reported value should use a prefix or scientific notation, but not both, unless you are comparing values. Scientific notation should usually use the base unit.

F.3 Quantitative Descriptions and Rules

Quantities that have *exact values* are either *defined* quantities (for example, 1 t is defined as exactly 1000 kg, and the SI prefix *kilo*, k, is exactly 1000), or quantities obtained by *counting* (for example, 32 people in a class or any coefficient in a balanced chemical equation). You can be certain about such quantities; there will be a small degree of uncertainty when counting very large numbers.

On the other hand, most quantities are measured by a person using some measuring instrument (for example, measuring the mass of a chemical using a balance). Since every instrument has its limitations and no one can perfectly measure a quantity, there is always some uncertainty about the number obtained. This uncertainty depends on the size of the sample measured, the particular instrument used, and the technological skill of the person doing the measurement.

Accuracy

Accuracy is an expression of how close an experimental value is to the accepted value. The comparison of the two values is often expressed as a percent difference. For example, the accuracy of a prediction based on some authority can be expressed as the absolute value of the difference divided by a predicted value and converted to a percent.

$$\% \text{ difference} = \frac{|\text{experimental value} - \text{predicted value}|}{|\text{predicted value}|} \times 100$$

This expression of accuracy is often used in the Evaluation section of investigation reports.

Precision

Accuracy is an expression of how close a value is to the accepted, expected, or predicted value, whereas ***precision is a measure of the reproducibility or consistency of a result*** (**Figure 27**). Accuracy is generally attributed to an error in the system (a *systematic error*); precision is associated with a *random error* of measurement. For example, if you used a balance without zeroing it, you might obtain measurements that have high precision (reproducibility) but low accuracy. The systematic error might be high (low accuracy), but the random error of the measurement is low (high precision).

Scientists define precision as the closeness of the agreement between independent measurements. We make the assumption that, if an instrument produces a certain decimal fraction (like a tenth of a unit), then all repeated measurements would be the same except for that last digit.

As long as an instrument is read correctly, for simplicity, we will assume that ***precision is the place value of the last measurable digit and is determined by the instrument.*** A mass of 17.13 g is more precise than 17.1 g. The precision is determined by the particular system or instrument used; for example, a centigram balance versus a decigram balance.

You may not know how uncertain the last measured digit is. On a centigram balance, the error of measurement in the last digit is usually ±0.01 g. Measurements such as 12.39 g, 12.40 g, and 12.41 g all have the same precision (hundredths), and may all be equally correct masses for the same object. The precision with which you read a thermometer might be ±0.2 °C (for example, 21.0 °C, 21.2 °C or 21.4 °C) and a ruler might be read to ±0.5 mm; you must decide, for example, whether to record 11.0 mm, 11.5 mm, or 12.0 mm.

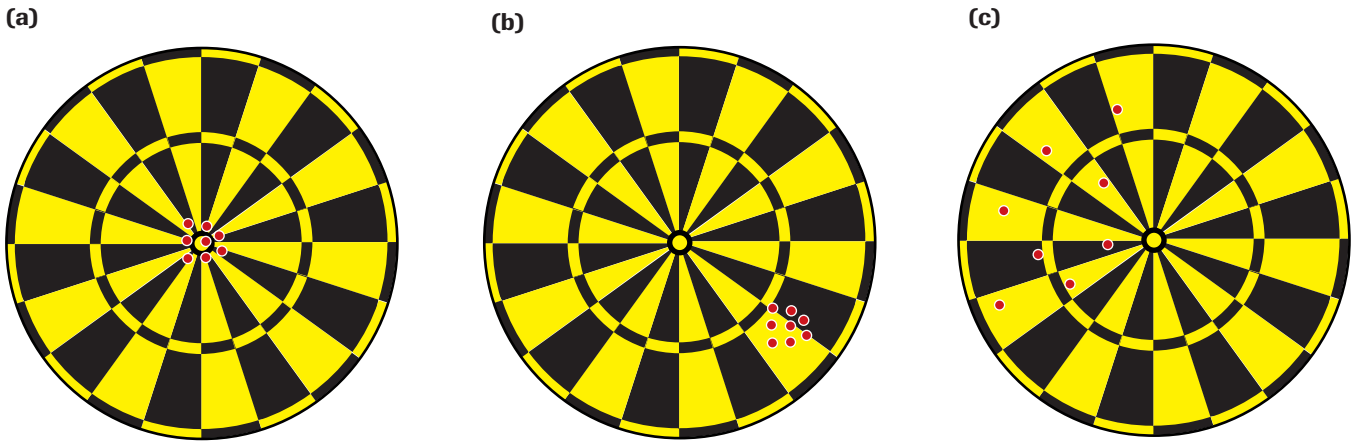

Figure 27
The positions of the darts in each of these figures are analogous to measured or calculated results in a laboratory setting. The results in **(a)** are precise and accurate, in **(b)** they are precise but not accurate, and in **(c)** they are neither precise nor accurate.

Precision Rule for Calculations

A result obtained by adding or subtracting measured values is rounded to the same precision (number of decimal places) as the least precise value used in the calculation. For example, 12.6 g + 2.07 g + 0.142 g totals to 14.812 g on your calculator. This value is rounded to one-tenth of a gram and reported as 14.8 g because the first measurement limits the precision of the final result to tenths of a gram and the rounding rule suggests leaving the 8 as is. The final result is reported to the least number of decimal places in the values added or subtracted.

Precision Rule for pH

The precision rule for pH and pOH is a special case, although the logic is consistent with that for values expressed in scientific notation. A hydrogen or hydronium ion concentration of 1.0×10^{-7} mol/L converts to a pH of 7.00. Just as the ^{7} is not counted as a significant digit when communicating the scientific notation value, the 7 is also not counted when communicating the pH value. Therefore, the rule is

- *The number of digits following the decimal point in a pH or pOH value is equal to the number of significant digits in the corresponding hydronium or hydroxide ion concentration.*

and

- *The number of significant digits in a hydronium or hydroxide ion concentration is equal to the number of digits following the decimal place in the corresponding pH or pOH.*

Certainty

How certain you are about a measurement depends on two factors—the precision of the instrument and the value of the measured quantity. More precise instruments give more certain values; for example, 15.215 °C as opposed to 15 °C. Consider two measurements with the same precision, 0.4 g and 12.8 g. If the balance used is precise to $\pm$0.2 g, the value 0.4 g could vary by as much as 50%. However, 0.2 g of 12.8 g is a variation of less than 2%. For both factors—precision of instrument and value of the measured quantity—the more digits in a measurement, the more certain you are about the measurement. We communicate the **certainty** of any measurement by the number of significant digits. In a measured or calculated value, ***significant digits are all those digits that are certain plus one estimated (uncertain) digit.*** Significant digits include all digits correctly reported from a measurement, except leading zeros. Leading zeros are the zeros at the beginning of a decimal fraction and are written only to locate the decimal point. For example, 6.20 mL (3 significant digits) has the same number of significant digits as 0.00620 L.

For each of the following measurements, the certainty (number of significant digits) is stated beside the measured or calculated value:

0.41 mL	a certainty of 2 significant digits
700 mol	a certainty of 3 significant digits
0.020 50 km	a certainty of 4 significant digits
2×10^{40} m	a certainty of 1 significant digit

Certainty Rule for Calculations

Significant digits are primarily used to determine the certainty of a result obtained from calculations using several measured values. ***A result obtained by multiplying or dividing measured values is rounded to the same certainty (number of significant digits) as the least certain value used in the calculation.*** For example, 0.024 89 mol $\times$ 6.94 g/mol is displayed as 0.1727366 g on a calculator. This is correctly reported as 0.173 g or 173 mg because the second value used (6.94) limits the final result to a certainty of three significant digits.

Research in the News

News media often quote the results of surveys, such as the percentage of people who would vote for a certain political party. What does it mean when we hear, "The results were Yes 52%, No 42%, Undecided 5%, with a margin of error of 3% 19 times out of 20"? The *margin of error* (3%) is usually calculated as the reciprocal of the square root of the sample size. A larger sample size would therefore produce a smaller margin of error. The *confidence level* (19 times out of 20) is like the precision. If the survey were repeated 20 times, the result would be within the percent error 19 times and 1 time it would be very different. The pollster, in effect, claims to be (accurate) within 3% of the "real" answer 19 out of 20 times.

Rounding

When completing calculations that involve more than one step, there are two rules that are used for answers in this textbook:

- *Never* round off partial answers in your calculator.
- *Always* round off when communicating partial answers on paper.

When chained calculations involve both multiplication/division and addition/subtraction, you may be required to store the partial answers in your calculator memory or to use the bracket function on your calculator.

Rounding Rule

Calculations are usually based on measurements (for example, in the Analysis section of a report). To report a calculated result correctly, follow this procedure. *Check the first digit following the digit that will be rounded. If this digit is less than 5, it and all following digits are discarded. If this digit is 5 or greater, it and all following digits are discarded, and the preceding digit is increased by one.*

F.4 Tables and Graphs

Both tables and graphs are used to summarize information and to illustrate patterns or relationships. Preparing tables and graphs requires some knowledge of accepted practice and some skill in designing the table or graph to best describe the information.

Tables

1. Write a descriptive title (**Table 7**).
2. The row or column with the manipulated variable usually precedes the row or column with the responding variable.
3. Label all rows and/or columns with a heading, including units in parentheses where necessary. Units are not usually written in the main body of the table.

Table 7 The Reaction of HCl(aq) with Zn(s)

Concentration of HCl(aq) (mol/L)	Time for reaction (s)
2.0	70
1.5	80
1.0	144
0.5	258

Graphs

1. Write a descriptive title on the graph and label the axes (**Figure 28**).
   - Label the horizontal (x) axis with the name of the manipulated variable and the vertical (y) axis with the name of the responding variable.
   - Include the unit in parentheses on each axis label, for example, "Time (s)."
2. Assign numbers to the scale on each axis.
   - As a general rule, the data points should be spread out so that at least one-half of the graph paper is used.

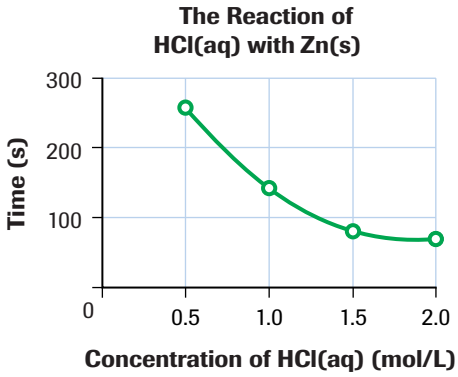

Figure 28
A sample graph

   - Choose a scale that is easy to read and has equal divisions. Each division (or square) must represent a small simple number of units of the variable; for example, 0.1, 0.2, 0.5, or 1.0.
   - It is not necessary to have the same scale on each axis or to start a scale at zero.
   - Do not label every division line on the axis. Scales on graphs are labelled in a way similar to the way scales on rulers are labelled.
3. Plot the data points.
   - Locate each data point by making a small dot in pencil. When all points are drawn and checked, draw an X over each point, or circle each point in ink.
   - Be suspicious of a data point that is obviously not part of the pattern. Double-check the location of such points, but do not eliminate the point from the graph if it does not align with the rest.
4. Draw the best-fitting line.
   - Using a sharp pencil, draw a line that best represents the trend shown by the collection of points. Do not force the line to go through each point. Uncertainty of experimental measurements may cause some of the points to be misaligned.

- If the collection of points appears to fall in a straight line, use a ruler to draw the line. Otherwise, draw a smooth curve that best represents the pattern of the points.
- Since the data points are in ink and the line is in pencil, it is easy to change the position of the line if your first curve does not fit the points to your satisfaction.

Using Your Graph

Although a graph is constructed using a limited number of measured values, the pattern may be used to extend the empirical information.

- *Interpolation* is used to find values between measured points on the graph.
- *Extrapolation* is used to find values beyond the measured points on a graph. A dotted line on a graph indicates an extrapolation.
- The scattering of points gives a visual indication of the uncertainty in the experiment. A point that is obviously not part of the pattern may require a remeasurement to check for an error or may indicate the influence of an unexpected variable.

F.5 Problem-Solving Methods

Definition of Terms

- **ratio**: comparison of two numbers; e.g., $\frac{2}{3}, \frac{1.5\text{ mol}}{1\text{ L}}$
- **proportion**: an equality of two ratios; e.g. $\frac{n_{NH_3}}{2.0\text{ L}} = \frac{1.5\text{ mol}}{1\text{ L}}$
- **conversion factor**: a specific type of ratio that is used to convert a quantity from one unit to another unit; e.g., $\frac{1000\text{ m}}{1\text{ km}}$ or $\frac{1\text{ km}}{1000\text{ m}}, \frac{40.00\text{ g}}{1\text{ mol}}$ or $\frac{1\text{ mol}}{40.00\text{ g}}$
- **formula**: a mathematical statement of a relationship using SI and IUPAC symbols and format; e.g., $m = nM$ (Note: m = n × M is not acceptable.)

Basic Problem-Solving Methods

A stoichiometry example is used to compare the three basic methods.

Copper metal is used to recover silver from a silver nitrate solution. Predict the mass of silver obtained from the complete reaction of 50.0 g of copper.

$$\underset{50.0\text{ g}}{Cu(s)} + 2\,AgNO_3(aq) \rightarrow \underset{m}{2\,Ag(s)} + Cu(NO_3)_2(aq)$$

63.55 g/mol — 107.87 g/mol

Proportion Method

$$\frac{n_{Cu}}{50.0\text{ g}} = \frac{1\text{ mol}}{63.55\text{ g}} \qquad n_{Cu} = 0.787\text{ mol}$$

$$\frac{n_{Ag}}{0.787\text{ mol}} = \frac{2}{1} \qquad n_{Ag} = 1.57\text{ mol}$$

$$\frac{m_{Ag}}{1.57\text{ mol}} = \frac{107.87\text{ g}}{1\text{ mol}} \qquad m_{Ag} = 170\text{ g}$$

Formula Method

$$n_{Cu} = \frac{m}{M} = \frac{50.0\text{ g}}{63.55\text{ g/mol}} = 0.787\text{ mol}$$

$$n_{Ag} = 0.787\text{ mol} \times \frac{n_{required}}{n_{given}} = 0.787\text{ mol} \times \frac{2}{1} = 1.57\text{ mol}$$

$$m_{Ag} = nM = 1.57\text{ mol} \times 107.87\text{ g/mol} = 170\text{ g}$$

Conversion Factor Methods

(a) Step method

$$n_{Cu} = 50.0\text{ g} \times \frac{1\text{ mol}}{63.55\text{ g}} = 0.787\text{ mol}$$

$$n_{Ag} = 0.787\text{ mol} \times \frac{2}{1} = 1.57\text{ mol}$$

$$m_{Ag} = 1.57\text{ mol} \times \frac{107.87\text{ g}}{1\text{ mol}} = 170\text{ g}$$

(b) Full method ("Factor Label")

$$m_{Ag} = 50.0\ \cancel{\text{g Cu}} \times \frac{1\ \cancel{\text{mol Cu}}}{63.55\ \cancel{\text{g Cu}}} \times \frac{2\text{ mol Ag}}{1\ \cancel{\text{mol Cu}}} \times \frac{107.87\text{ g Ag}}{1\ \cancel{\text{mol Ag}}} = 170\text{ g}$$